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A LABORATORY MANUAL

TO ACCOMPANY

GODFREY'S ELEMENTARY CHEMISTRY

LONGMANS, GREEN, AND CO.

Educ T 229.10.432

TABLE OF EQUIVALENTS

1 inch = 2.54 centimeters.
1 foot = .3048 of a meter.
1 mile = 1.6093 kilometers.
1 centimeter = .3937 of an inch.
1 meter = 1.0936 yards.
1 cubic inch = 16.387 cubic centimeters.
1 cubic foot = 28.317 cubic decimeters.
1 cubic centimeter = .061 of a cubic inch.

1 square centimeter = .155 of a square inch.
1 square meter = 1.196 square yards.
1 grain, Troy = .0648 of a gram.
1 ounce, Avoirdupois = 28.35 grams.
1 ounce, Troy = 31.104 grams.
1 gram = .03527 of an ounce, Avoirdupois.
1 gram = .03215 of an ounce,
oy.
= 15.432 grains, Troy.
gram = 2.2046 pounds,
oirdupois.

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D MEASURES

1 pound = 5760 grains.
1 ounce = 480 grains.

IGHT
ound . . . . lb.
undred-weight . . . . . . cwt.
on . . . . T.

IGHT
ms = 1 ounce . . . . oz.
ces = 1 pound . . . . lb.

RE
. . . . ft.
. . . . yd.
. . . . rd.
. . . . mi.
feet, or 633,360 inches.

ABLE
. . . . cm. = $\frac{1}{100}$ M.
. . . . dm. = $\frac{1}{10}$ M.
. . . . M.
. . . . Dm. = 10 M.
. . . . Hm. = 100 M.
. . . . Km. = 1000 M.
. . . . Mm. = 10000 M.

METRIC TABLE OF WEIGHTS

10 milligrams (mg.) = 1 centigram = cg. = $\frac{1}{100}$ gm.
10 centigrams = 1 decigram = dg. = $\frac{1}{10}$ gm.
10 decigrams = 1 gram = gm.

1000 grams = 1 kilogram = kg.
The cubic centimeter, cu. cm., is the common measure of capacity.
1000 cu. cm. = 1 liter = L.

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A LABORATORY MANUAL

ELEMENTARY CHEMISTRY

BY HOLLIS GODFREY

With many illustrations. Crown 8vo. Pp. xiv—456.

The book consists of 33 chapters dealing with those portions of Elementary Inorganic and Organic Chemistry which the author has considered of most value to the student, with a final chapter on Chemical Arithmetic (Part I); brief Epitomes of the 33 descriptive chapters for purposes of review (Part II); and Questions drawn from the individual chapters (Part III).

A LABORATORY MANUAL

TO ACCOMPANY

GODFREY'S ELEMENTARY CHEMISTRY

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Author of "The Health of the City," etc., etc.

LONGMANS, GREEN, AND CO.
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1915

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PREFACE

THIS laboratory manual is intended to accompany Godfrey's Elementary Chemistry. It contains eighty experiments, of which the first sixty-four are designed to occupy two laboratory periods a week for thirty-two weeks and correspond to the text book work of the thirty-three descriptive chapters of the Chemistry. There are no experiments given for Chapter XIII. on Equations. This omission is due to the author's belief that at this point, where the first introduction of the subject of equations occurs, the whole of the laboratory time should be given over to practice in equation writing. The first sixty-four experiments are chiefly qualitative. The last sixteen experiments are partly quantitative. These latter experiments furnish material for about thirty-two additional periods of laboratory work, afford some opportunity for choice and offer to the pupil with an excess of time a wider range of subjects. Exercise 80 on the bending of glass is placed, for convenience of reference, at the end of the book.

A few points characteristic of this manual should receive some mention. A brief note precedes each experiment, referring the laboratory work directly to the occurrences of every-day life. Repetition is an essential part of learning with most of us, and the impressions derived from the statements of the text gain in strength by the reënforcement of the manual. Careful detailed directions are given in every case. Questions intended to bring out the salient points of each experiment end each exercise. Important topics have been italicized. A continual effort has been made to advise the use of simple, inexpensive apparatus and to require only simple manipulation. Out of the eighty exercises given here sixty-eight can be performed with the simple apparatus in the hands of every pupil, such as flasks, beakers, and test tubes. Twelve experiments require slightly more expensive apparatus, such as burettes and condensers. These may be made lecture-table experiments. Illustrations of some of the more complicated pieces of apparatus have been included in the list of plates in order to show pupils the method of adjusting the apparatus mentioned in the text. A full list of apparatus and of chemicals will be found in the appendix.

No claim of originality in the exercises chosen for this manual is made. To a marked degree chemical science has already determined what experiments are practical for use in the ordinary school laboratory. The sources whence these experiments have been drawn are so various, and this manual has been in the process of construction for so many years, that detailed acknowledgment would be impossible. One acknowledgment, however, should be made. From the Boston Teachers of Chemistry, with whom the author has met month after month and year after year, have come many suggestions of great value.

The purpose of this laboratory course is similar to the purpose of Godfrey's Elementary Chemistry — constantly to bring before the pupil's mind the connection between the laboratory, the text-book and the affairs of every-day life. The author recognizes the possibility of varying conditions existent in different schools and welcomes all criticism and advice which will aid in the difficult task of joining the practical affairs of every-day life to the limitations of the ordinary school laboratory.

Besides being read by several readers to whom the author expressed his acknowledgments in the preface to the chemistry proper, the manuscript of this manual has been read by Miss L. C. Smith, instructor in chemistry in the Brookline High School, and by Miss Florence M. Homer, the author's present assistant.

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A LABORATORY MANUAL

LABORATORY DIRECTIONS

1. Keep your locker in order. Clean desk and sink before you leave the laboratory.

2. Provide a wiping cloth and a laboratory coat or apron.

3. Before heating, wipe all glass and porcelain dry. If any liquid spills as you proceed, stop heating until the dish is wiped dry.

4. Take up all chemicals with weighing paper.

5. Use no matches except the safety matches provided. Return them to place after use.

6. Leave no solids or matches in the sinks.

7. Leave no apparatus out of your locker except by special permission.

8. Do not place bottle stoppers on the desk. Hold them in your hand or place them on paper. Return all reagents and chemicals to place immediately after use. Never move carelessly when carrying solutions.

9. Heat test tubes in as nearly a horizontal position as possible. Heating them in a perpendicular position will cause the liquids within to spill over.

10. Crucibles, ignition tubes, and test tubes are heated over the bare flame. Beakers, flasks, and evaporating dishes must always be heated on asbestos or on wire gauze above a flame. Wide-mouth bottles and gas bottles made of heavy glass must never be heated over a flame.

NOTE-BOOK WRITING

Two things should be obtained from any course in science: ability to see and ability to express what is seen. Every experiment should give greater power to understand the wonders of natural phenomena. Every set of notes should give greater power of expression. A systematic and orderly method of entering the notes aids the student and lightens the labors of the teacher. Every set of notes should be properly headed and dated. A division of the main body into three parts, under the general topics of Method, Results, and Answers to Questions, while not essential to the proper understanding of an experiment, makes references to former experiments and corrections of the work accomplished easier. A model set of notes follows. Your notes should have the same general form as that shown here.

EXERCISE I

(Date here.) 8/12/09. (Name here) JOHN SMITH.

Physical and Chemical Change
Effect of heating in air

1. *Method.* Heated copper in a crucible and observed its condition before and after heating. Heated sulphur in a test tube, observing the changes which took place during the process.

2. *Results.* The copper before heating was bright and metallic. After heating it looked dull and black. This dark appearance seemed to be permanent. It did not rub off or change after the metal cooled. Apparently a new substance was formed.

The sulphur changed to a dark red-brown liquid, then to a gas which changed back to a solid again in the upper part of the tube and left yellow solid sulphur. This looked just the same as the solid sulphur we had at first. The sulphur at the end seemed to be the same substance with which we started.

B

1. The sulphur passed through a physical change.
2. The copper passed through a chemical change.
3. The copper became another substance by heating.

In entering your own notes you need not enter the words " Method," " Results," or " Answers to Questions," as the numerals 1 and 2 and the letter B indicate these things.

Leave one line between each division of the work and write on one side of the paper only. Commence each experiment on a fresh sheet.

Place name and date on the first line, not on the upper part of sheet. The letter " M " on a corrected experiment means that you are to bring your note-book to the instructor. The letter " F " means " Finish and submit."

There are two things which you must know before you begin any experiment.

I. *What am I going to study in this experiment?*

II. *How am I going to perform this experiment?*

Mechanical following of directions is almost worthless. Each experiment has a definite end and a definite method for accomplishing that end. It is your business to understand both. Every experiment in this book states its purpose in the heading. Notes beneath the heading discuss any points in relation to the experiment that you should know. Then comes the apparatus needed. Directions for performing the laboratory work follow, while in Part B are questions on the results obtained.

The first thing to be done in every experiment is to take out your note-book and enter the heading and date. Enter the notes as the work progresses.

The last thing to be done in every experiment is to look up the list of individual apparatus needed for the next exercise. This list is entered near the beginning of each experiment. The general apparatus such as balances and burners, common to all students, need not be considered. See if you have everything necessary. If not, report the apparatus needed to the instructor or enter the missing articles under your name and locker number in the " Apparatus Supply Book."

EXERCISES

The notes on this first exercise should serve as a model for your future notes. Perform this experiment, read the model notes, and then try to rewrite them in your own language.

Exercise 1. Physical and Chemical Change. Effect of Heating in Air. — In this experiment we are to heat copper and sulphur in air. One of these substances undergoes a *physical change*, one, a *chemical change*. The problem before you is to determine by the use of your eyes which substance has undergone some *radical change* in structure, and which remains *unchanged* in structure.

Note. — Examine your bunsen burner before commencing this experiment. You will notice that it has a hole at its base covered by a movable slide. Close this slide and light the burner. You see that the flame is yellow and gives out light like the flame of an ordinary illuminating gas burner in a house. Now turn the slide so that the base hole is open and note the change. The flame is now blue in color and does not give illumination as it did before. It is similar to the flame of a gas cooking stove. If you look at it carefully you will see that there is a blue cone in the center of the flame. Later on we shall have a chance to know more about this flame, but for the present it is sufficient to say three things.

I. The flame that gives light is called a *luminous flame*. It is used in bending glass, but for little else in this course. It is produced when the base-hole of the burner is closed.

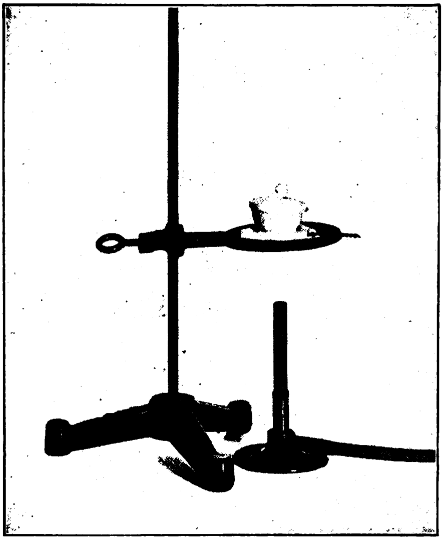
II. The blue flame which gives little light is called a *non-luminous flame*. This is the flame used in most of the work performed in this course. It gives out more heat than the luminous flame. Its hottest point is probably just above the tip of the blue cone. It is produced when the base-hole of the burner is kept open.

III. Even when the base-hole is open the flame sometimes becomes suddenly luminous and smoky. This change is due to what is called the "snapping back" of the flame. The gas instead of burning at the top of the burner burns also at the base. When this occurs, turn off the gas at once and let the burner cool or cool it with water. Relight when it is cold.

Apparatus — Crucible, crucible triangle, and test tube.

1. Describe some powdered copper as fully as you can. Heat it

in a clean crucible for five minutes, noting all the changes that take place.



1. Heating a Crucible. — The crucible rests on a pipe clay triangle placed on the ring stand. It is exposed to the direct heat of the flame. Flasks, beakers, and evaporating dishes, on the other hand, are never exposed to the direct flame. They are heated on asbestos or on wire gauze. The bottom of the crucible should be but a slight distance above the top of the blue cone of the flame.

2. Do the same with sulphur, heating it in a test tube instead of in a crucible.

B

1. Which substance apparently passed through a *physical change* by heating?

2. Which substance apparently passed through a *chemical change* by heating?

3. Did either one seem to become a new substance? If so, which?

Exercise 2. Chemical Processes — Precipitation, Decantation, Filtration and Evaporation. — In this experiment we are to use four of the operations commonly carried on to facilitate chemical work. In *precipitation* a solid and liquid are formed from the union of two liquids. In *decantation* a solid and a liquid are partially separated by pouring off the upper liquid (called the supernatant liquid) from the settled solid. You are really decanting when you pour tea or coffee from tea leaves or coffee grounds. In *filtration* a solid is separated quite thoroughly from a liquid by means of porous paper which allows the liquid to pass through while the solid remains behind. A cook uses filtration whenever fruits are strained for jelly. *Evaporation* turns a liquid into a gas which passes off into the atmosphere. We evaporate water whenever we boil it. Solids which are present in the liquid

2. PRECIPITATION. — The beaker on the right contains a clear solution of one compound. That on the left contains the solution of another and a different one. These two solutions are represented in the experiment which precedes, by lead nitrate and salt respectively. The center beaker shows the moment of precipitation when the two solutions added together produce a solid precipitate.

are left behind in this process. In consequence solids may be separated from liquids by evaporation.

Apparatus — Beaker, stirring rod, funnel, flask, filter paper, evaporating dish.

Precipitation. — Place about three teaspoonfuls of a solution made from a compound of lead (called lead nitrate) in a test tube and add three times as much water. Pour a similar amount of a solution of common salt into a test tube and dilute it with a corresponding amount of water. Now pour the two solutions into a beaker and note the white solid that has been formed from the two liquids. The solid is called a *precipitate.* The action is called *precipitation.*

Decantation. — Let the mixture remain in the beaker until the solid sinks to the bottom. When this has occurred pour the *supernatant liquid* carefully off from the solid, letting a small amount run into an evaporating dish. Pour the rest into the filter as described below. This method of separating a liquid from a solid is called *decantation.* While waiting for the solid to settle proceed with the next part of the experiment.

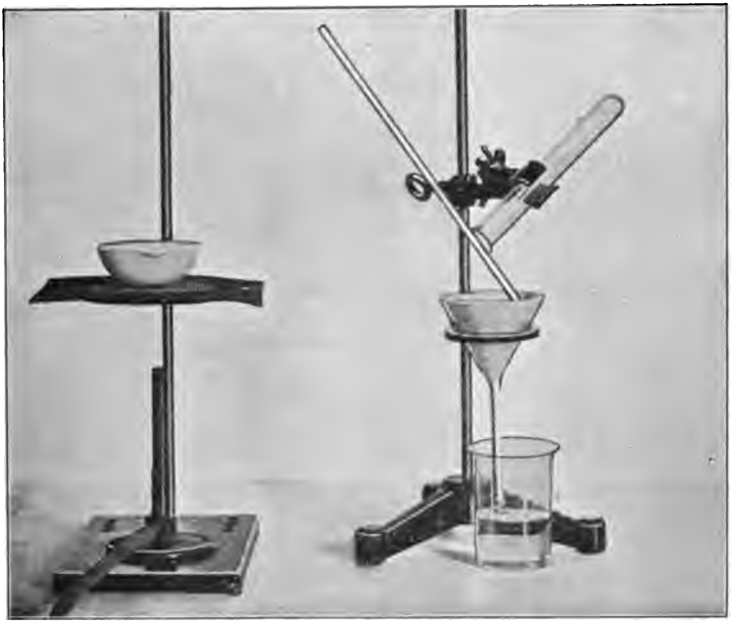
Evaporation and Filtration. — When the solid has partially separated from the liquid and you have poured a small amount of the supernatant liquid into an evaporating dish, place the dish on asbestos and light the burner. Have the tip of the blue cone just below the dish. Allow the liquid to *evaporate* (boil away) to dryness, turning off the gas as soon as the liquid disappears. While this is going on do the work on filtration. Fold a filter paper and place it in the funnel as shown in the illustration. Moisten the paper slightly to hold it in place. Then place your funnel in a ring stand and pour the substance remaining in the beaker through the filter paper into the flask. Pour down a glass rod and be careful not to spill.

The solid *precipitate* stays behind. The liquid *filtrate* goes through. This action is called *filtration.*

B

Write a brief description of the four processes.

Exercise 3. Weight and Volume. Weight of a Crucible in Grams. Capacity of a Crucible in Cubic Centimeters. — We make practically all our purchases by *weighing* and *measuring.* We buy a pound of sugar and a quart of milk. We are to weigh and measure in the metric system here. The following rules should be followed in all weighing: Always place the substance to be weighed on the left pan of the balance. Never place anything on the balance which will injure the pans. When you are in doubt whether or not possible injury to the balance may result from any substance, weigh it in a previously weighed crucible or a beaker. Keep your small weights, tenths and hundredths of a gram, on a small piece of paper while



3. Evaporation and Filtration. — On the left a solution is being evaporated, that is, the liquid of the solution is changing to a gas under the influence of the heat produced by the lighted gas which comes from the gas burner below. (This type of burner is called a Bunsen burner.) To keep the evaporating dish from cracking it is placed upon an asbestos centered wire gauze during the operation. On the right a liquid holding a solid (like water holding sand) is being poured from a test tube, down a glass rod and into a glass funnel lined with filter paper. Filter paper is porous and allows liquids to pass through while it holds solids back. Such a separation is called filtration. The liquid which passes through is called a filtrate. The solid which remains is called a residue when precipitation has not taken place.

you are weighing. One invariable rule must be remembered in all weighing done in the laboratory. It follows:

When you weigh, take your note-book to the balance with you and immediately enter the weight obtained.

Apparatus — Crucible.

Part I. Weigh a crucible without its cover on the scales. Enter the weight in your notes, writing the amount in *grams* and *tenths of grams.* For instance, should the crucible weigh thirteen grams, six decigrams, and four centigrams, enter this as 13.64 *gms* *not* as 13 grams, 6 decigrams, 4 centigrams. Record as follows:

Weight of crucible =

Part II. A glass whose capacity has been measured and marked by lines in such a way that we may tell at a glance the amount of liquid inside of it is called a *graduate.* A common size of graduate measures twenty-five cubic centimeters. A *liter graduate* will measure one thousand cubic centimeters. Graduates offer an easy means of directly calculating the volume of any liquid, or of computing the capacity of any vessel. We are to determine the capacity of a crucible here.

4. Weighing. — Balances on ring stand. Gram weights and forceps below. Small weights on paper are fractional parts of a gram. Large weights in block run from 1 gram to 100 grams.

Fill your crucible up to the brim with water and then carefully pour the water into the graduate. Read the number of *cubic centimeters* it contains and record as follows:

Capacity of crucible in cubic centimeters (cubic centimeters are commonly abbreviated to cc.) =

Remembering that one liter contains one thousand cubic centimeters, determine what fractional part of a liter your crucible contains and enter as follows:

Capacity of crucible in fractional parts of a liter =

(In recording use decimals always.)

Part III. One cubic centimeter of water weighs, practically speaking, one gram. Refer back to Part II. in which you determined the capacity of the crucible in cubic centimeters and state the number of *grams of water* in crucible when it is full.

Number of grams of water held in crucible =

Exercise 4. Density, Amount of Matter per Unit of Volume. — The density of any body is the *amount of matter* (*or weight*) *per unit of volume.* A piece of iron, for example, that weighed 71 grams and had a volume of 10 cubic centimeters would weigh 7.1 grams per cubic centimeter and would have a density of $\frac{7.1}{1}$. In this case we consider it as *the number of grams to each cubic centimeter.* We find the density of the wood we are to use by dividing the weight, measured in grams, by the volume, measured in cubic centimeters.

Apparatus — **20 cm. rule.**

Take a wooden block, of the kind used to support apparatus and determine its volume in cubic centimeters. Then find its weight in grams. Enter each number in the same fashion as before, making no entries on this page, but entering each measurement in your notebook as soon as taken. Do all your numerical work in your notebook.

Data

Length = Breadth = Thickness = Volume = Weight =

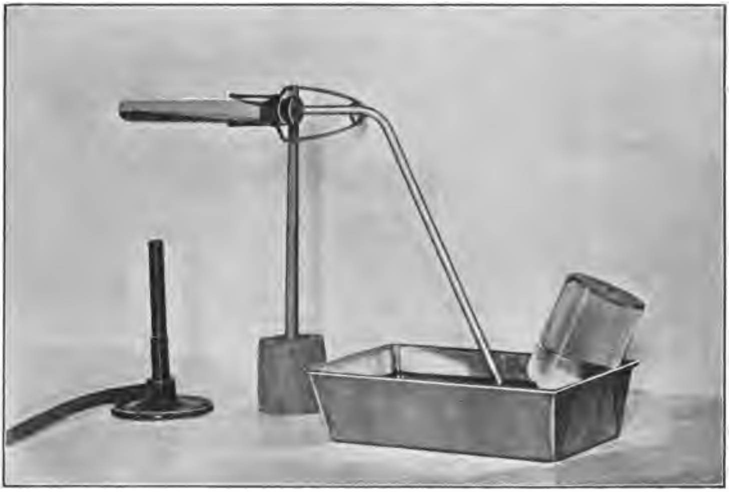
Density =

Exercise 5. Preparation of Oxygen from a Compound of Oxygen and Mercury called Mercuric Oxide. — We use this gas, *oxygen,* which we are to prepare here, every day of our life for breathing, for making fires, for preparing various necessary compounds.

In this experiment you are to make oxygen from a compound of mercury and oxygen. From this compound the gas, oxygen, may be evolved by the aid of heat. This is the way in which Priestley, the discoverer of oxygen, formed the gas. You are also to make a *flame test.* This test is commonly used. By it we may distinguish one gas from another, noting whether the gas puts out a flame, supports a flame with difficulty, increases a flame or allows a flame to burn smoothly. In this special case you are to consider the action of the gas with regard to a glowing splinter, not a brightly lighted one.

NOTE. — The expression "flame test" is used commonly for a test for metals in their solutions, which we shall try later. In its use in the experiments on gases we refer to the action of these gases on flames as related above.

As this experiment is the first in which you collect a gas, a word about the process is given here. The *collection of a gas over water* is



5. PREPARATION OF OXYGEN. — Either mercuric oxide or a mixture of potassium chlorate and manganese dioxide is being heated in a hard glass tube. The gas passes out through a bent delivery tube into the receiving bottle placed in a pneumatic trough (here a cake tin). Methods for bending glass tubes are discussed in the last exercises in this manual.

possible only when the gas is insoluble. When this is the case a gas will displace water from a bottle placed like the bottle in the illustration. Before beginning the experiment proper, set up the apparatus collecting the gas. To do this fill the pneumatic trough two-thirds full of water. (A pneumatic trough is a dish used for the collection of gases. A cake tin when it serves this purpose becomes a pneumatic trough.) Then fill a wide-mouth bottle full of water, close it with a watch glass and place it mouth down under the surface of the water. The water does not run out of the bottle. The reason it does not is because the air outside, pressing down upon the surface of the water in the pneumatic trough, holds up the water within the bottle which is not subject to similar downward pressure. Another point to be considered with regard to the collection of gas is found in the fact that air, present in the tubes which convey the gas, must be removed before we can test a gas. In consequence the first bubbles of gas evolved should always be allowed to go free.

Apparatus — Two ignition tubes, delivery tubes, wide-mouth bottle, pneumatic trough, watch glass.

1. Heat some mercuric oxide in an ignition tube connected with a delivery tube so arranged as to lead any escaping gas into a small test tube or an ignition tube, filled with water and inverted in a pneumatic trough as the bottle was inverted. The thumb may be used to close a tube of this size instead of a watch glass. Gas is evolved during the heating. When the receiving tube is full of gas take the delivery tube out of the water and let it cool. Leave the receiver full of gas in the trough until it is needed. If the delivery tube is not removed from the pneumatic trough, cold water may be sucked back into the hot apparatus. This is likely to break the glass. Remember this point in future experiments.

2. Light a piece of slow-match. Take a tube of collected gas and turn it quickly, mouth up. Plunge the slow-match into the ignition tube now full of the gas, *oxygen*, that came from the mercury powder. Note the flame and product as before.

3. Note appearance of the *oxygen* as it stands in the bottle.

4. Hold the lighted piece of slow match in a bottle of air. Note the size and color of the *flame* and describe the product of the burning. The phrase "*product of burning*" means in this case the visible product left, as the ash formed. There are products of burning which are invisible gases. At the end place any mercury powder left in the tube in the bottle labeled "Mercury Powder Residues." *Do not let any mercury get into the sink.* Be careful here, as in all experiments where mercury is used, *to keep the mercury away from gold rings.*

B

Does the oxygen formed look like air?

Does the slow-match burn in it more or less brightly than in air?

Does the flame of the slow-match look the same in this gas as in air?

Do you think the solid product left by the burning slow-match is the same when it is burned in air as when it is burned in this gas?

Exercise 6. Study of Oxygen. Its Preparation from Potassium Chlorate and Manganese Dioxide. — One point should always be remembered in collecting gases. They have a great tendency to escape into the atmosphere. After a bottle of gas has been collected, a glass plate, or better still, a watch glass, should always be slipped over the mouth of the bottle, only to be removed when the gas is tested. Gases heavier than air are kept in bottles having the mouth up. Gases lighter than air are kept in bottles with the mouth down. Oxygen is very nearly the same weight as air and it is tested in a bottle kept mouth up for convenience sake. On account of lack of time you may be unable to write the notes on this experiment during the laboratory hour. If you cannot do this, enter the essential data in the laboratory, finish and submit the notes the next laboratory hour. Make this a general rule when an exercise takes so much time as to make the writing of notes during the hour impossible.

In the following experiment you are to form oxygen from potassium chlorate, a substance rich in oxygen, with the assistance of manganese dioxide, a substance which does not give off any oxygen, but which does help the white crystals of potassium chlorate to set free the gas. After obtaining oxygen you are to study it with relation to its action with a flame and with relation to the compounds which it forms.

Apparatus — Ignition tube, delivery tube, bottle, pneumatic trough, deflagrating spoon.

Mix a small amount of potassium chlorate with about half as much manganese dioxide in a mortar; *very carefully.* (If pounded too hard *they may explode.*) Then place the mixture in an ignition tube and heat gently, collecting the gas as in the previous experiment. When placing the mixture in an ignition tube let it lie along the lower side of the tube so as to give the *oxygen* evolved a chance to come off through the whole length of the tube. Be careful to keep the mouth of the tube slightly raised. Fill and refill the bottle as needed, testing as follows: Into the first bottle full of gas, plunge a glowing splinter. In the second, place a bit of burning sulphur on a deflagrating spoon. Into the third, lower a bit of hot charcoal on a deflagrating spoon. Into the fourth, lower a little heated *red* phosphorus on a deflagrating spoon. Into the fifth, lower a piece of iron watch spring which has been dipped in sulphur and the sulphur lighted. Watch for the burning of the iron here, not for the burning of the sulphur.

Be sure that your lighted splinter, phosphorus and sulphur are extinguished by plunging them into water when you finish.

B

Do substances burn better in air or in oxygen?

Do the products formed by burning substances in this gas seem to be the same as those formed in air?

What does the answer to the preceding question show concerning the presence of oxygen in air?

Exercise 7. First Study of an Acid. Action of Metals on Sulphuric Acid as Compared with their Action on Water. — This experiment takes considerable time to perform, but it can be done in a laboratory hour if you will take two or three short cuts which do not affect the value of the experiment. These short cuts follow. If you cannot write up your notes during the hour, complete them immediately afterwards and bring them in next hour.

Commence Part II as soon as you have placed all your metals in the liquids in Part I. Use two funnels and filter only a few cubic centimeters of each liquid into the test tubes. Filtering and evaporating both take time and, when only qualitative results are desired, much time can be saved by using only a small amount of liquid. With care you can safely evaporate these liquids in test tubes held nearly horizontal, but you must be careful not to spill any liquid and to have the outside of the tube clean. Theoretically all evaporating should be done in shallow dishes like evaporating dishes which keep a large surface of liquid exposed from which the gases may rise, but in the laboratory we can usually evaporate small amounts to dryness in test tubes.

In this experiment you are to take two typical *metals*, zinc and iron, and place them, first in an *acid* (in this case sulphuric acid), and afterwards in *water*, comparing the action of the *metal* on the *water* with the action of the *metal* on the *acid*. There is a marked difference in the chemical action of each. You are also to consider a problem of *solubility*. You know that water dissolves sugar but not sand. You are to see if you can find any evidence of solubility in this experiment, and especially to notice if any new substance whose general characteristics you can recognize later is produced.

NOTE.—When heating anything in a test tube always hold the mouth of the tube away from you. Take care not to get the tube very hot.

Apparatus — Test tubes.

Make the following experiments in test tubes, pouring in a sufficient quantity of dilute acid to about one-sixth fill the tubes. Take four test tubes. Place pieces of zinc in two, a very small pinch of iron filings in two. Use one piece of metal in each tube.

Now pour
- Sulphuric acid on zinc in tube 1.
- Sulphuric acid on iron in tube 2.
- Water on zinc in tube 3.
- Water on iron in tube 4.

Note especially whether or not the metals appear to have dissolved. See if any *gas* comes off. Save the test tubes and their contents.

B

How does the action between the sulphuric acid and the metals differ from the action with water?

What evidence was there that the phenomena were due to the acid and not to the metal?

Compare the solubility of the metals in water with their solubility in acid.

Did you see anything that looked like a gas?

Part II. Compounds Formed by the Action of Metals upon Sulphuric Acid — Filter the liquids left in the test tubes after the preceding action from any *solids* which remain, filtering through different papers into different test tubes, and evaporating a small amount of each liquid in test tubes. Stop heating as soon as any *white solid* appears on the side of the test tubes. Carefully describe any solid residues.

NOTE. — When sulphuric acid reacts with a metal, the metallic compound formed is ordinarily a *sulphate.*

B

How does the residue left from the test tubes containing sulphuric acid and a metal differ from the residue left from those containing water and a metal?

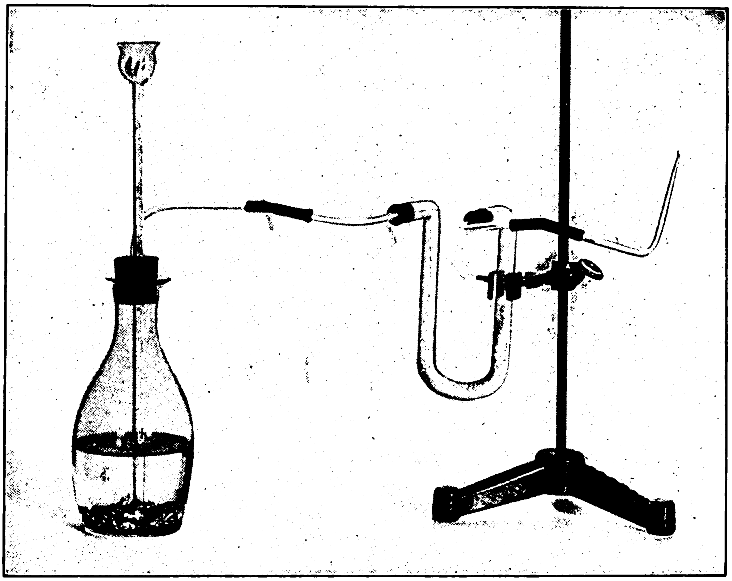
Name the new substances formed by the action of the sulphuric acid on the metals by reference to the note above.

Exercise 8. Study of Hydrogen. — In this experiment we are to examine a phenomenon which seems a wonderful thing — the product formed when hydrogen, a light inflammable gas, burns in oxygen or in air. Part I of this exercise, the preparation of *hydrogen,* is ordinarily sufficient to occupy the time of a whole period. If there is an extra laboratory hour the class can do Part II and burn hydrogen. The *burning of hydrogen* is an experiment which requires ample time and should not be hurried. It is often advisable for the class to perform the first part of this experiment and for the instructor to show as a lecture-table experiment both the collection of hydrogen by upward displacement and the burning of the gas.

In all experiments with hydrogen, the *greatest care* must be taken. Hydrogen itself is not explosive, but hydrogen mixed with oxygen is *explosive. Never have a flame near a hydrogen generator* and always add the acid to the metal slowly. Here we are to form *hydrogen* by performing one part of Exercise 7 on a larger scale, that is, we produce the gas by the action of more sulphuric acid on a larger quantity of zinc. We are particularly to consider some of the properties of hydrogen. Note the word "generator" just used. Any apparatus in which gas is produced is called a generator. Gases may be generated in a tube, a flask, or a gas bottle such as is shown here. When we refer to a generator such as is used here we shall mean a gas bottle generator.

Apparatus — Gas bottle, two-hole stopper, delivery tube, thistle tube, pneumatic trough.

Set up a gas bottle with two-hole stopper, delivery tube, and thistle tube as shown in the illustration, and connect with a pneumatic trough as in the experiment with oxygen. For this part of the experiment you need not connect the generator with a drying tube. Place three or four pieces of granulated zinc in the bottle, cover them with water, put on the stopper, and pour as much dilute sulphuric acid as you have used of water through the thistle tube. Let the chemicals act for a couple of minutes to clear the generator from air and then collect two bottles of the gas over a pneumatic trough. (Remember the point already mentioned, — the generator must be freed from air whenever you generate a gas. Every vessel open to the atmosphere and not filled with anything else is full of air, and this air must be driven out before you can get the gas or test its properties.) Light the gas in



6. Preparation of Hydrogen. — The gas is generated in a thick glass bottle fitted with a thistle tube and a delivery tube. The thistle tube dips below the surface of the liquid. The delivery tube reaches only to the bottom of the stopper. The gas may either be collected by displacement over water, as oxygen was collected, or by upward dry displacement as here. In this latter case the calcium chloride *drying tube* removes water present in the gas.

one bottle, holding the bottle mouth down. Light the match and hold the bottle *away from the generator*. *Be sure* that the burning hydrogen in a lighted bottle of the gas is wholly *extinguished* before bringing it near the generator again.

Take the second bottle full of gas and, after leaving it standing upright and uncovered on the table for a couple of minutes, apply a lighted match as before.

At the end of the experiment pour the acid out of the generator. Wash the zinc by decantation and place it in the bottle of zinc residues. Be careful not to let any zinc fall into the sink.

B

Does hydrogen burn?

Does it allow a splinter to burn in it?

What is the flame test for oxygen?

What is the flame test for hydrogen?

What became of the hydrogen that filled the bottle left mouth upward?

In what ways is this gas like or unlike oxygen?

Part II. The Burning of Hydrogen and the Product formed by its Burning.— Set up the generator as shown in the illustration, replacing the delivery tube by a right bend, inserting a drying tube and attaching a tip to the bend by means of a rubber connector. Collect the gas by allowing it to pass upward through a tip into a test tube held mouth down. The gas rises in the test tube and drives out the air. This method of collecting gas is called collecting by upward dry displacement. Test with a match.

Allow the gas to run until, on lighting a test tube full of the gas, no report is heard. Then ask the instructor to light the tip. Hold a cold dry test tube near the flame. Describe results.

B

What do you think appeared on the walls of the test tubes?

What kind of a substance forms when anything is burned in air?

What two elements does this product probably contain?

What is true of the weight of all gases that can be collected by upward displacement?

What would be the weight of a gas that you could collect by downward displacement?

Exercise 9. Acid, Base and Salt. Action of Acids, Bases, and Salts on Indicators. — Indicators are chemical substances which show by a change in color whether a solution in which they are present is acid or alkaline. In this experiment you are to try four kinds of indicators with acids, bases, and salts. After it, you should be able to tell at a glance whether any substance is an acid or alkali by using one of these indicators. Phenolphthalein, one of the best indicators we possess, cannot, however, be used with a carbonate. Two common acids are to be tested here, sulphuric acid and vinegar, two common bases, ammonia and caustic soda (caustic soda is commonly used in the household in place of lye), two common salts, table salt and sodium sulphate (the salt to be prepared in the next experiment).

Apparatus — Test tubes, red and blue litmus paper, turmeric paper.

Copy the following table into your note-book before coming into class, using it in place of the ordinary statement of method and results.

Substance to be Tested	Color from Methyl Orange	Color from Phenol-phthalein	Effect on Red and Blue Litmus Paper	Effect on Turmeric Paper
Sulphuric acid				
Vinegar				
Ammonia (Ammonium hydroxide)				
Caustic soda (sodium hydroxide)				
Salt (sodium chloride) ...				
Sodium sulphate				

Prepare in six test tubes solutions of sulphuric acid, vinegar, ammonia, sodium hydroxide, salt, and sodium sulphate. The solutions should be approximately 1 to 10. (That is, they should consist of 1 part of acid, base, or salt to 10 parts of water. *Pour the acid on the water in mixing.* Never pour water on acid.)

Put a few drops of methyl orange solution into each test tube and carefully note the color, recording it in the table in your book.

Repeat the whole experiment with fresh solutions, using this time a few drops of phenolphthalein. Carefully note results. Test each solution by dipping a stirring rod into the solutions and then dropping a single drop on small strips of red and blue litmus paper. Note changes and record results. Do the same with turmeric paper.

Part II. Pour the ammonia solution into the hydrochloric acid, which has been treated with the phenolphthalein and note any color changes. If no change occurs add more ammonia until a change in colors just occurs and record the result.

B

Which of these substances are acid, which are alkaline, which are neutral?

How can you neutralize an acid? That is, how can one remove its acid qualities?

By what indicators can one tell whether a solution is acid or alkali?

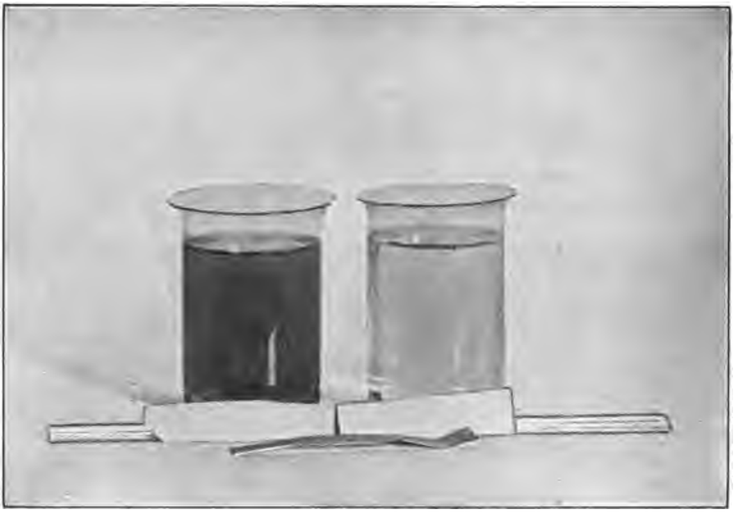
At the end of this experiment take home four slips of red litmus and four slips of blue litmus paper. Test with them any fruits in the house, and any household materials, such as starch, lye, sugar, household ammonia, soap, etc., which you think might give results. Report in writing at the end of the notes on this experiment.

Exercise 10. Preparation of a Normal Salt, Sodium Sulphate.—The majority of the substances that we meet in every-day life are neither acid nor base, but *neutral*. Not all neutral substances by any means are *salts*, but many of those we meet in the laboratory are. We are to form one salt here. In this experiment you are to neutralize an acid with a base, producing, as a result of the action, a neutral salt and water. Exact neutralization, such a neutralization that an indicator would show that a solution was neither acid nor base, is impracticable in the laboratory. By slight additions of acid and base we bring the solution to a point where it is just over the edge of neutralization. We obtain this point practically when a single drop of base turns an indicator one color and a single drop of acid turns it the other. In the use of the indicator employed here, phenolphthalein, you have reached the proper point when a drop of sulphuric acid added to the solution turns the indicator white and a drop of sodium hydroxide turns it pink.

After the action is completed the neutral salt obtained is left in solution in water. Evaporating the water leaves the salt.

The method used in this experiment to produce a *sulphate* (the salt ordinarily produced by the action of sulphuric acid) is to be used in several later experiments. It is, therefore, important that you should thoroughly master the process given here in order that you may be able to use it easily hereafter.

A sulphate is commonly prepared by adding sulphuric acid to a metal or to the oxide or hydroxide of a metal. One of the two former methods was illustrated by the action of sulphuric acid on zinc. The other may be illustrated by the simple action of sulphuric acid on zinc oxide.



7. NEUTRALIZATION. — Hydrochloric acid and sodium hydroxide treated with the indicator phenolphthalein. The *acid* on the right is colorless. The *base* on the left is pink. Small pieces of litmus and turmeric paper rest against the beakers, together with the stirring rod, which should be used in all neutralization experiments.

These two ways of forming zinc sulphate, shown in the two following equations, may be performed after completing Exercise 10.

Zinc + hydrogen sulphate = zinc sulphate + hydrogen.

Zinc oxide + hydrogen sulphate = zinc sulphate + water (hydrogen oxide).

The action of a metallic hydroxide upon an acid is illustrated here.

Apparatus — Flask, beaker, evaporating dish, medicine dropper, stirring rod.

Pour into a flask 100 cc. of water, and 10 cc. of sulphuric acid. Pour into a beaker 100 cc. of water and add 10 cc. of sodium hydroxide solution. This sodium hydroxide solution has been made by dissolving solid sodium hydroxide in water. Pour enough of the dilute sulphuric acid solution just made into your evaporating dish to about one quarter fill it, and add six drops of phenolphthalein solution. Then very carefully and slowly pour in the alkali solution, stirring constantly until the color changes. If your evaporating dish gets too full, empty out half of the solution and continue. After the color has changed, reverse the action by aid of a little acid, added, drop by drop, by means of a medicine dropper. Continue this action until one or two drops of sodium hydroxide will restore the color or one or two drops of the sulphuric acid will remove it. The solution is then practically *neutral.* Test it with litmus and turmeric paper. Finally, pour a few cubic centimeters of the solution into a test tube and evaporate to dryness. Examine the sodium sulphate formed.

NOTE. — In most of the exercises, which follow, the action can be expressed in the form of a chemical equation. Such equations may be written in words or in chemical formulæ. So the first equation given above may be written:

Zinc + hydrogen sulphate = zinc sulphate + hydrogen.

or

$$Zn + H_2SO_4 = ZnSO_4 + H_2$$

Before the introduction of the study of chemical symbols and formulæ, however, the equations should be written in words. After that time they should be written in symbols.

B

Complete this equation. Sodium hydroxide + hydrogen sulphate =

What tests do you know by which we can tell whether or not a solution is alkaline?

What tests do you know by which we can tell an acid?

What tests do you know by which we can tell a neutral?

Complete this general equation. Acid + base =

Exercise 11. Study of Water. Determination of the Freezing Point of Water. Comparison of the Temperatures of Water at the Boiling

Point and of Steam. Comparison of the Freezing and Boiling Points of a Solution of Salt. — Note: For this experiment jars filled with ice alone, and with ice and salt, should be prepared. The spaces between the ice should be filled with water. Water in the ice of the ice-chest, the water of the faucet, and the steam of the kettle is the commonest subject of every-day household experiment. We are to discuss here some of the temperature conditions of water and also the temperature conditions of that commonest of solutions, salt in water. The results may throw some light on the reason why we salt the water when we want potatoes to cook rapidly.

In this experiment we are to use a Centigrade thermometer for the first time. All thermometers depend for their action upon the expansion of some liquid or gas. The liquid metal mercury is generally used for this purpose. Thermometer scales are calculated by obtaining the expansion or increase of volume of the liquid or gas chosen, when it is heated from the freezing to the boiling point of water. The Fahrenheit scale, which is the one we commonly use, has 32° for its freezing point and 212° for its boiling point. There are 180° between the two. The Centigrade scale, which is the one used in all scientific work, has 0° for its freezing point and 100° for its boiling point. There are 100° between the two.

Apparatus — Flask, asbestos, burner.

Take a thermometer and place it in one of the jars containing ice water. Leave it there while you are preparing to do the rest of the experiment. Then take the reading, record it, and proceed.

Fill a flask half full of water and boil the water. When vigorous boiling has begun, take the temperature of the liquid and of the steam inside the flask. Take the reading when the bulb is wholly in water or wholly in steam. Hold the thermometer by means of a twist of wire. Keep it immersed for about one minute. Be sure that it does not touch the bottom or side of the flask. Take away the flame for two or three minutes, then pour about a heaping teaspoonful of salt into the hot water in the flask. Find the boiling point of this solution. Record the result and compare with the boiling point of water. Then place your thermometer in one of the jars containing ice, salt, and water, and record results.

B

What is the relation between the temperature of boiling water and the steam it gives off?

What is the relation between the freezing and boiling points on the Centigrade and Fahrenheit thermometer scales?

What is the relation between the boiling and freezing point of water, and of a solution of salt in water according to your results in this experiment?

What relation do these last facts have to the freezing of ice-cream? to the boiling of potatoes?

Exercise 12. Study of Solution. Comparison of the Effect of Hot and Cold Water on the Solution of a Solid. Crystallization. — We perform the first part of this experiment at the table whenever we put a considerable amount of sugar into solution in hot tea and then let the liquid cool. Every time we do that we are studying the effect of hot and cold water on solutions. We see water crystals whenever we watch snow crystallize on the window pane. Many substances in passing from a liquid or gaseous state into that of a solid assume a definite shape, solidifying into regular forms bounded by plane surfaces. This change is called crystallization and crystals are formed by the process. Crystals may be obtained by cooling a saturated solution, by slowly evaporating a solution, or by cooling a melted solid or vapor. We shall try the first two methods here.

8. A CRYSTAL.

9. ANOTHER CRYSTALLINE FORMATION.

Apparatus — Beaker, asbestos, stirring rod, funnel, evaporating dish, watch glass.

Saturate a boiling solution with sodium nitrate, the sodium salt of nitric acid. In order to do this take a beaker and add 15 grams of sodium nitrate to 20 cc. of water at the temperature of the room. Stir until dissolved. Add 15 grams more and note if *complete solution* seems to take place. After stirring for a short time, put the solution on an asbestos and heat to boiling. Stir until all the solid is dissolved. The solution should then be practically saturated, that is, it should have taken all the sodium nitrate into solution that it will hold. While heating the solution, prepare a funnel and filter paper. When the solid has mostly disappeared, stop heating and, as rapidly as possible, filter most of the hot liquid into your evaporating dish. While pouring hold the hot beaker by means of a wiping cloth. Filter a few drops of the liquid into a watch glass and, if *crystals* do not form at once, leave it until the next laboratory hour. Examine the evaporating dish and watch glass and note results.

B

Which dissolves the greater quantity of most soluble solids, hot or cold water?

What are three ways of making crystals?

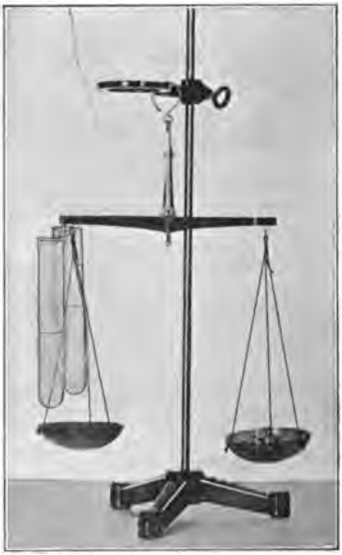
Can you note any difference in type between the solutions formed in this experiment and the solution of zinc sulphate made from zinc and sulphuric acid when we were studying hydrogen? In this case we started with sodium nitrate and had the same thing when we got through. How was it in the case of the zinc and sulphuric acid?

Exercise 13. An Illustration of the Law of the Conservation of Matter. — To prove the great laws on which our modern scientific theories and processes are based has required many years of careful investigation by the most expert and learned scientists. We can illustrate these laws only roughly with the apparatus at our command, but we can at least show some indication of the way in which they act. In the following experiment we are endeavoring to show that the two factors which enter into a reaction weigh the same as the two differently composed factors which come out of it, or, in general terms, we are trying to illustrate the law of the conservation of matter by showing the relation between the weight of two substances both in solution, and the weight of the substances produced by their reaction when one is a solid and the other a liquid.

Apparatus — Two test tubes.

Arrange your apparatus according to the illustration. One tube contains a dilute solution of barium chloride; (made by using 1 gm of crystallized barium chloride to 10 cc of water). The other tube contains dilute sulphuric acid; (1 cc. acid to 10 cc. of water). Shake the tube containing the solid barium chloride until all has dissolved, and then counterpoise the tubes on the balance with sand or weights.

Pour the contents of the first tube into the second without spilling. Then divide the mixture equally between the two tubes. Hang them exactly as before. Remember that a very slight change in the point from which the tubes are hung will throw the balance off and seriously affect the experiment.



10. Apparatus for Ex. 13. — Two test tubes are connected by a loop of copper wire which runs through the hook which supports the pans. Great care should be taken that the tubes are hung from the same point in both weighings and that they have the same amount of liquid in each before and after the precipitation.

B

Has the weight increased or decreased?

Write the reaction for the chemical change. How does the sum of the weights of the products after the reaction compare with the sum of the weights of the factors before the reaction?

This experiment illustrates the general law of the conservation of matter. State the law and show how it applies not only in this case but in preceding ones.

Exercise 14. An Illustration of the Law of Definite Proportions by Weight. — This experiment, illustrating roughly the law of definite proportions by weight, can be done satisfactorily by the class only when there is a sufficient supply of burettes. Where time is short, the solutions may well be prepared in the required quantity a sufficient time in advance to save the trouble which comes from irregular mixing. In any case, owing to the length of time which it takes to perform this exercise, this can well be made a lecture-table experiment. The chief difficulty in this experiment is in the variation of the strength of acid and alkali solutions. The ratio between the two should not be more than $\frac{1.5}{1}$. If the quantities indicated do not give this, change to stronger or weaker solutions as desired.

Apparatus — Flask, gas bottle, beaker, filter paper, stirring rod.

If solutions of hydrochloric acid and sodium hydroxide are prepared, use the stock solutions. If not, make a dilute solution of hydrochloric acid in your flask by adding 4 cc. of concentrated acid to 200 cc. of water. Take 10 cc. of a strong solution of sodium hydroxide and pour it into 200 cc. of water in your gas bottle, stirring thoroughly.

After these solutions have become thoroughly mixed, each and every cubic centimeter of them should contain the same *amount by weight* of the acid or alkali used. Therefore, if we use 20 cc. in one case and use 10 cc. in the following case, we shall use *twice* as much by weight in case 1 as in case 2.

Enter the following data in your note-book.

A No. cc. of acid solution used.	B No. cc. of alkali solution used.	Ratio of $\frac{(A)}{(B)}$
1st Trial.		
2d Trial.		
3d Trial.		
4th Trial.		

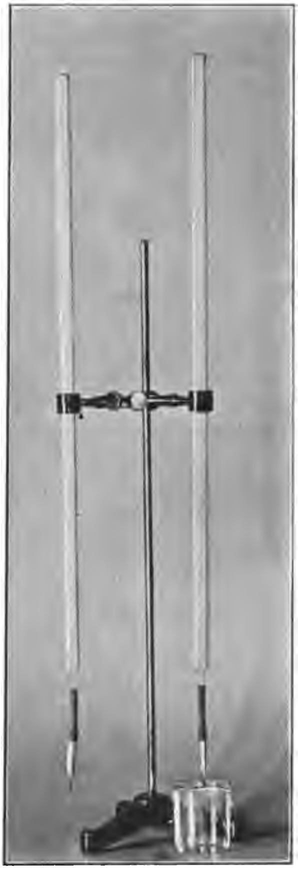
A sample set of data would read as follows:

	Acid (A)	Base (B)	$\frac{A}{B}$
1st Trial	45 − 25 = 20 cc.	48 − 8 = 40 cc	$\frac{20}{40} = \frac{.5}{1}$ $\frac{.5}{1}$

NOTE. — You will note that the burettes are marked like a graduate. If you commence to draw off acid when the liquid stands at 25 and end at 45 you have, of course, drawn off 20 cc.

Take a pair of burettes from the table and see that they are clean. Call one burette A, the other burette B. Burettes with ground glass stoppers are for acid or neutral solutions only. Never put an alkaline solution in a burette of this kind. Note carefully which one is for the acid and which one is for the alkali. Fill burette A with hydrochloric acid solution and burette B with sodium hydroxide solution. Draw off a few cc. from each burette. Make sure that the tube below the burette stop is filled with the liquid. Note exactly the height of the liquid in each burette, reading from the lower line. *Note and record the height of the liquid in both burettes.* Draw into a clean beaker 20 cc. of the acid solution and add to it a few drops of phenolphthalein solution. Place a clean piece of filter paper under the beaker.

Now from burette B draw sodium hydroxide into the solution until the solution is a delicate pink color which is permanent. Remove the color by adding acid from burette A, and add first one and then the other until one drop gives a *permanent change.* (Add solutions *cautiously* and at the last drop by drop, *mixing thoroughly* with a stirring rod after each addition.) Read the burettes at the end and record the height of each liquid.



11. BURETTES ON A BURETTE STAND. — In using burettes it is well to put a piece of filter paper under the beaker in which neutralization is taking place. Make sure that no air bubbles are left in the burette with the rubber tube by turning the tip obliquely upward and running a few cc. through it.

Make no effort to get even numbers of cubic centimeters. Take whatever quantities neutralize and obtain your data by subtracting the upper reading from the lower.

Enter your data in the table shown. Enter the exact number of cc. and tenths of cc. of acid and alkali under A and B. Get the data required in the third column by dividing the number of cc. of hydrochloric acid required by the number of cc. of sodium hydroxide used.

Continue the experiment by using different amounts of acid solutions, 15 cc., 30 cc., and 35 cc., using a sufficient amount of alkali to bring about neutralization. Record the data in the table.

B

From this experiment make a statement as to the proportions in which substances unite.

Give the full statement of the law we are endeavoring to illustrate.

Exercise 15. A Series of Flame Tests. Preparation of Apparatus for Making Hydrochloric Acid. — In the following experiment we are to use the *flame test* for a metal. We are using this phrase " flame test " in connection with the solution of a metallic compound in a different sense than when we used it with a gas. Different metals in solutions when heated in the non-luminous flame, or by other methods, give characteristic colors to such a flame, or show certain illuminated lines when their light is passed through a prism. We test metals in the laboratory in a simple fashion by putting their compounds into solution and heating a few drops of this solution on a platinum wire. We use platinum for this purpose instead of any other metal because this metal in this form does not affect the color of the flame. Sodium, the metal present in common salt, is one that we are to use here, and sodium is one element which gives the characteristic color to the flame produced. In the second part of the experiment we are to make ready for the preparation of hydrochloric acid gas which is to be studied in the next experiment.

Apparatus — Flask, two-hole stopper, thistle tube, delivery tube, watch glass, cake tin, litmus paper, filter paper, two wide-mouth bottles, gas bottles, asbestos.

Test a solution of salt in the flame by means of a platinum wire. Record the result of this test. In a similar fashion test nitrates of lithium, potassium, calcium, strontium and barium. Hold the wire in the flame until the color disappears on each test. At the end mix together a small amount of the solutions of sodium and potassium. Observe the color. Then look at the flame through a cobalt glass. The light passing through this glass gives but one of the colors. See if you can think what happened to the other.

Now proceed to prepare your apparatus for next time as follows:

Place about 10 gms. of salt on a paper and pour into your flask 40 cc. of sulphuric acid made up by adding 25 cc. of concentrated sulphuric to 20 cc. of water.

Remember always to pour acid into water.

Place the acid in a *flask*, not in a gas bottle, stopper the flask with a two-hole stopper fitted with a thistle tube and a delivery tube, and set it away on the closet floor, where it will not spill over; see also that all apparatus entered above is collected.

B

What is the flame test for sodium? for lithium? for potassium? for strontium? for barium? for calcium?

What was the use of the cobalt glass?

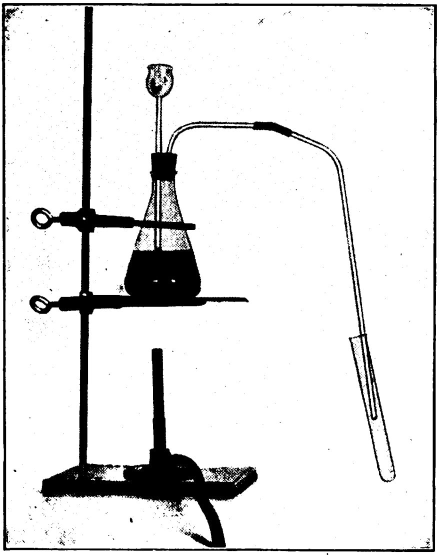
Exercise 16. Study of Hydrochloric Acid Gas. — The hydrochloric acid made in this experiment is the much used " muriatic acid " of commerce. Do this experiment *under a hood* or draft closet. In all use of the hood *keep the windows closed. Do not inhale this gas or let it get outside of the hood.* Where hood space is deficient this experiment can be performed by one or two groups and demonstrated to the rest of the class. This statement is true of all the experiments which must be performed under a hood.

Apparatus — Same as in last experiment.

Place the apparatus, prepared last time, on asbestos, on a ringstand under the hood in the fashion shown in the illustration. Place the lower ring high enough to allow a sufficient flame below. (As a general rule in heating a flask place it so that it is well above the tip of the blue flame.) Place there also two empty wide-mouth bottles, a test tube, and a gas bottle full of water. Add the salt to the acid, heat gently, and collect two bottles of the gas by downward displacement. Note the color of the gas and test the full bottles immediately with the flame test *for a gas*. Cover the mouths of the bottles with watch glasses as soon as the bottles are filled. Then, while testing the gas in the bottles, let the generator run into your gas bottle full of water for about five minutes.

Test with moist strips of red and blue litmus paper held over the mouth of the bottle.

Test for *solubility* by inverting a test tube full of the gas in a pneumatic trough two-thirds full of water. Note any change of the water level within the tube.



12. Preparation of Hydrochloric Acid and Chlorine. — An Erlenmeyer flask containing sulphuric acid and salt is placed on asbestos as shown above, and heated gently to produce hydrochloric acid. The gas is collected by downward displacement in the test tube. Chlorine is prepared in a similar apparatus by heating hydrochloric acid and manganese dioxide in the flask. Both these gases should be kept *under a hood.*

As a characteristic test for *hydrochloric acid,* run a little of the gas into a test tube containing silver nitrate (a solution made by adding one part of silver nitrate solution to ten parts of water). Test with litmus paper the liquid into which the gas has been flowing. At the end of the hour open the windows, close the door, and pour the liquid remaining in the generator into the sink. Wash out the apparatus and collect everything ready for the next experiment, which uses the same apparatus.

B

What are the four chief characteristics of this gas?
You added salt to sulphuric acid. Write the equation.
What is the characteristic test for hydrochloric acid gas?

NOTE. — In preparation for the next experiment bring some brightly colored pieces of printed cotton cloth.

Exercise 17. Study of Chlorine. — This experiment illustrates the preparation of chlorine, the gas which bleaches. Although of great value in the manufacture of paper and textiles, it is chlorine in washing powders, which is largely responsible for the way in which laundry work bleaches and ruins colored fabrics. You are to consider the process of preparation and some of the properties of the gas, especially its bleaching action. *Do this under a hood. Do not inhale the gas.*

In this experiment chlorine is prepared by the action of manganese dioxide upon hydrochloric acid. The oxygen of the manganese dioxide combines with the hydrogen of the hydrochloric acid, and the chlorine of the acid in part unites with the manganese and in part goes free. It is this free chlorine that we are to test.

Apparatus — Flask, two hole rubber stopper, thistle and delivery tubes, two wide-mouth bottles, asbestos, burner, bits of variously colored cotton cloth.

Place your apparatus under the hood, set up as it was in Experiment 16, when you made hydrochloric acid. Put into the flask about as much manganese dioxide as you used of salt in the last experiment, stopper the flask, and pour about 20 grams of strong hydrochloric acid through the thistle tube. Heat gently and collect the gas formed by downward displacement in two bottles. Let the generator run into a gas bottle full of water for two minutes. Finally stop heating. *Do not inhale the gas* or let it get outside the hood. Test by putting a lighted match into the gas. Note the *color*, test with *wet* litmus paper and *wet* cloth. Test the gas for solubility by inverting a test tube full of the gas in a pneumatic trough two-thirds full of water and leaving it there for a minute. Take the bottle full of water into which the gas has been running and test it with litmus. Place it on your desk, leave it until the next day and test it again. Note the results. Fill another bottle with chlorine solution and save it for Ex. 20. At the end of the hour pour out any liquid remaining in the generator and wash out the bottle, having the windows open and the door closed to avoid the spread of the gas.

B

Describe the properties of chlorine as fully as possible, weight as compared with air, color, solubility, bleaching action.

The manganese dioxide sets free oxygen. What action does the oxygen have upon the hydrochloric acid?

What happened to the solution of chlorine in water after it was placed in the sunlight, and why did it happen?

Exercise 18. Common Salt (Sodium Chloride) Produced by the Neutralization of Sodium Hydroxide with Hydrochloric Acid. — In this experiment we are to make *salt*, like the salt you used at breakfast this morning. In preparing salt in this way we are also illustrating two important chemical truths: one, that chlorides are ordinarily formed from hydrochloric acid, and the other that an acid neutralized by a base produces a salt. We have already seen this latter truth in our production of sodium sulphate. We shall practically duplicate that experiment here. Remember that, chemically speaking, " a salt " means a member of the group of salts of which our common salt, sodium chloride, is merely a type. In preparation for this experiment read carefully the directions and notes given under Exercise 10, " Preparation of Sodium Sulphate."

Apparatus — Beaker, flask, evaporating dish, stirring rod, litmus, and turmeric paper, test tube.

Neutralize 10 cc. of sodium hydroxide with hydrochloric acid exactly as you neutralized with sulphuric acid. You are to use in this case 10 cc. of the dilute hydrochloric acid on the desk in 100 cc. of water. Take 10 cc. of the sodium hydroxide solution on the desk and add 100 cc. of water. Neutralize carefully, using phenolphthalein as an indicator. After *neutralization* is complete, evaporate a few cc. of the salt solution to dryness. Try the flame test with a part of the solution by means of the platinum wire.

B

Write the equation which takes place.
Tell some of the uses of common salt.
Where does our ordinary salt come from?
What element does the flame test show to be present?

Exercise 19. Iodine and the Starch Test for Iodine. — This experiment illustrates certain properties of iodine, the ease with which it turns to a *vapor*, the way in which it may be tested with *starch*, and the way in which it *sublimes* or forms a *sublimate*. Sublimation is the process by which a substance changes directly from a solid to a gas or from a gas to a solid without apparently melting or passing through the liquid state. If you ever buy a bottle of solid iodine you will very likely find on the label "resublimed iodine." You should understand after this experiment what that means. Iodine is valuable as a counter-irritant in medicine, but the strength of the solution to be used for such a purpose should always be determined by a physician, as concentrated solutions are dangerous. Later on we shall study starch and you should remember that the starch test for iodine is also the iodine test for starch.

Apparatus — Test tubes, evaporating dish, asbestos.

Weigh out about a gram of starch and powder it in your mortar. After powdering, place it in your evaporating dish, add a few drops of cold water, and mix thoroughly. Then add 40 cc. of water and bring to boiling. When boiled you will have a starch paste. While the paste is heating take a dry test tube, drop in it a few crystals (not more than three or four) of potassium iodide, a little manganese dioxide and cover the mixture with sulphuric acid. Iodine is formed. Heat *carefully* in the flame. *Do not inhale* the vapor or *continue* the action after the vapor begins to come off. Note any change in appearance. Watch the upper end of the tube. When the starch has boiled dip a piece of filter paper into it and hold the paper at the mouth of the test tube. Take a small portion of the starch paste and add water. Test by adding a little iodine dissolved in water and a little more dissolved in alcohol. Compare the relative solubility of the element in these two liquids. Note the changes which occur and record.

PART II. — At the close of the experiment, *if time permits,* add a little sulphuric acid to potassium iodide. *Do this under the hood.* Study the compounds formed and note the reference note in the next experiment.

B

How is iodine prepared ?
What are some of the characteristic properties of iodine?
How may we test for iodine?
How could we test for starch?
What is the color of iodine vapor?

Exercise 20. — Comparative Test for Bromine and Iodine. — Bromine is chiefly known to most of us by the bromides of medicine and the bromide papers of photography. In this experiment we are to use a test by which we may compare a compound of bromine and one of iodine. We are to use as the reagent in this case a liquid known as carbon bisulphide. This substance is extremely *inflammable* and *no flame* should be lighted in the laboratory while the experiment is going on. BE CAREFUL. BROMINE VAPOR IS DANGEROUS.

Apparatus — Test tubes.

Prepare solutions of potassium iodide and potassium bromide by dissolving a small quantity of each compound in a test tube two-thirds full of water. To each tube add a little carbon bisulphide, and a little chlorine solution, shake the tubes and note results obtained. This is a means for comparing bromides and iodides.

Make bromine by a precisely similar action to that in which you made iodine *with the exception* that you use *potassium bromide instead of potassium iodide.* Note any change in appearance. Watch the upper end of the tube. Heat to *boiling only* and *do not inhale* the vapor. *Do this under the hood.* Pour the liquid down a sink at the end of the experiment, holding the tube away from you and running a considerable stream of water down the sink. *If time permits,* make hydrobromic acid by a like substitution, using potassium bromide and sulphuric acid instead of potassium iodide and sulphuric acid

Do this under the hood. Study the compounds formed. Pour the liquid down a sink at the close of the experiment, holding the tube away from you and running a stream of water down the sink.

NOTE — Hydrobromic acid is formed here, as hydriodic acid was formed in the preceding experiment. These actions are by no means simple, however. Both acids decompose readily. Bromine and iodine are formed. Sulphur dioxide is produced from the sulphuric acid and other compounds may occur.

B

Give some of the characteristics of bromine.

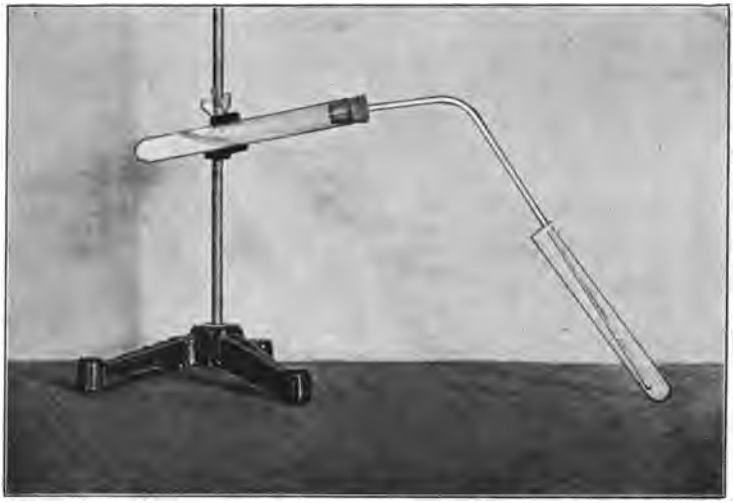
Describe the method by which you could distinguish an iodide from a bromide by the use of carbon bisulphide.

Compare the preparation of hydrobromic and hydriodic acids.

Exercise 21. Study of Carbon Dioxide Gas. — In this experiment we are to form *carbon dioxide gas*, the gas produced in all ordinary burning, the food of the plants. We really make carbon dioxide every time we light the wood of a match or build a fire in the stove. In order to study the properties of this gas satisfactorily, however, we must take some means which will enable us to separate it out and collect it by itself. Such a means is at hand in the evolution of carbon dioxide gas, when a member of the group of salts known as the *carbonates* is acted upon by an acid. Carbon dioxide is produced in seidlitz powders when the solution of one of the powders which contains an acid principle is added to the solution of another which contains a carbonate. We could use almost any carbonate and almost any acid to secure some quantity of the gas, but convenient materials to use are ordinary *powdered marble* (calcium carbonate) and *hydrochloric acid.* When you try to write the equation representing the action of this experiment remember that chlorides are regularly produced by the addition of hydrochloric acid. The test for the presence of carbon dioxide gas is made by means of lime water (calcium hydroxide). This is the ordinary lime water of the druggist. Simply note the change produced in making the test without considering the chemical action involved. We shall consider this test at length in a succeeding experiment. The action which takes place between the calcium carbonate and the hydrochloric acid illustrates a second method of preparing a *chloride*. Note that, while powdered marble serves well to produce the small quantity of carbon dioxide gas used here, marble chips, solid pieces of marble, give better results when a continuous flow of the gas is required.

Apparatus — Test tube with stopper, delivery tube, wide-mouth bottle.

Set up a generator like that shown in the illustration. Place marble in it, judging the amount by the quantity shown. Pour some water on the marble and add an equal quantity of hydrochloric acid. Collect the gas by downward displacement. Test the gas with lime water (calcium hydroxide) and a lighted match. Test the residue in the flask with a platinum wire in the flame in the same fashion as that by which you tested salt. Test the lime water precipitate with hydrochloric acid.



13. Making Carbon Dioxide. — Carbon dioxide is generated when marble and hydrochloric acid are mixed together. The action proceeds without the aid of heat. The gas is collected by downward displacement.

B

Does this gas support combustion?

By what test can you recognize this gas?

Write the equation for this action?
What is left in the generator?
What becomes of the chlorine in the hydrochloric acid?
What did the flame test show?

Exercise 22. Carbon Dioxide in the Room. Carbon Dioxide as Produced in the Body. The Burning of Wood. — In this experiment we are to consider the presence of carbon dioxide in the laboratory, in the body, and in the gases produced from the burning of wood.

Apparatus — Short piece of glass tubing, test tubes, watch glass.

Pour a small quantity of lime water into a watch glass and let it stand uncovered on your desk until the end of the hour. If no change occurs, leave it until after the close of recitations and examine it then. Take a clean piece of glass tubing, wash it carefully, and half fill a clean test tube with a clear solution of calcium hydroxide (lime water). (In all use of lime water care should be taken to have the solution clear.) Describe any change. Blow through the glass tube into the test tube until the lime water changes noticeably. Then hold a burning match inside of a test tube for a moment, remove it and, closing the mouth of the tube with your thumb, shake up the gas, left by the burning match, with lime water. Describe any change. If no change comes at first, try the experiment again. Compare all the changes which have occured.

B

The chief part of the wood that burns is carbon. What was formed by the burning match? What was found on the watch glass?

In the light of this experiment what must be going on in the body?

Exercise 23. Properties of Marble. — In studying the properties of this solid substance we are to consider solubility, appearance, action of acids and weight as compared to water. Water serves as the general standard of comparison for solids and liquids. We commonly compare substances by saying that they are heavier or lighter than water, that alcohol is lighter than water, granite heavier. In a similar way hydrogen and air serve as standards for gases. · The carbonate of calcium which we are to study here in the form of marble, occurs also as limestone and coral. In all its forms it will be found to have the same reactions. (Calcium, the metal of calcium carbonate, in many ways resembles sodium. A carbonate radical contains a carbon atom and three oxygen atoms and is written CO_3. Remember that a radical is a group of atoms acting like a single atom.)

Apparatus — Test tube, watch glass, litmus and turmeric paper.

1. Drop a little powdered marble into a test tube, add water, shake thoroughly, boil and then allow the marble to settle completely on to the bottom. Then very carefully, without stirring up the marble, by means of a medicine dropper, pour one drop of the liquid on your watch glass, which must have previously been cleaned. Let the drop dry without heating, and then examine the glass for a residue. Remember when we have evaporated solutions of solids that a considerable amount of solid matter has generally been left behind. Notice the quantity of the residue if any appears. Shake up the marble in the test tube and test with methyl orange. Test a little marble in two other test tubes with hydrochloric and sulphuric acids. Try the flame test for calcium.

B

Give all the properties of marble that you have discovered.

Is marble acid or alkaline?

If alkaline, what should be produced when we add an acid to it?

What is the action of hydrochloric and sulphuric acid on marble?

Bring some washing soda, some baking soda, and some tooth powder from home and test each with hydrochloric acid as you tested marble.

What kind of compounds appear to be present in these substances?

Exercise 24. Lime, Calcium Oxide. The Residue left from Heated Calcium Carbonate. — Lime is chiefly known as the white powder that forms mortar with sand and hair. An enormous amount of lime is used, however, in many manufacturing processes where it serves as an alkali. No other alkalies have the wide use of calcium oxide, unslaked lime, and of calcium hydroxide, slaked lime.

Lime, calcium oxide, is the residue left after heating calcium carbonate. It is generally obtained from limestone, but it is more convenient here to obtain it from powdered marble. *Quicklime* is calcium oxide, CaO. If we add water to quicklime we get *slaked lime,* calcium·hydroxide, $Ca(OH)_2$. CaO

$$\begin{array}{ll} & \mathrm{CaO} \\ & \underline{\mathrm{OH_2}} \quad (\mathrm{H_2O}\ \text{reversed}) \\ & \mathrm{Ca(OH)_2} \end{array}$$

A solution of calcium hydroxide in water is called *lime water.*

Apparatus — Covered crucible.

1. Balance a little marble, held in a covered crucible, with sand. Have the crucible one-third full. Heat at the hottest part of the flame for twenty minutes or more with crucible partly covered. While this is heating, get litmus paper and fill a pneumatic trough. At the end of that time put the cover closely on the crucible and weigh the whole as soon as cool. Test the residue in the trough for solubility and for alkalinity. The substance left after ignition is quicklime. Water added to this forms slaked lime. The solution formed when slaked lime dissolves in water is called lime water.

B

Is the crucible heavier or lighter after ignition?

What is the cause of the change?

Is the residue soluble?

Is it acid or alkaline?

What other substance that you have used, does this resemble in its action on litmus paper?

NOTE. No experiments have been written for the laboratory work which would naturally correspond to the work of Chapter XIII. in the text-book, as it is the belief of the author that the laboratory periods during this first introduction of equation writing should be given over to the practice of equations. From this point on every simple equation should be written in symbols and entered in the notes with the questions.

Exercise 25. Study of Nitrogen Gas. — The *nitrogen* we are to make here is the ordinary nitrogen of the air, the gas which holds in check the fires of the earth. When we want a considerable quantity of nitrogen it is simpler to make it by the combination of *ammonium chloride* and *sodium nitrite*, the method used here, than to separate it from the air by the use of phosphorus, a method explained later. The nitrogen obtained is produced from both the chemicals used and, when the action is carried to completion, the total nitrogen present may be obtained.

Apparatus — Flask, two-hole stopper, thistle and delivery tubes, two wide-mouth bottles, asbestos, pneumatic trough.

Place about equal quantities of ammonium chloride and sodium nitrite in a generator and about one-third fill the flask with water.

Set up the generator in the fashion shown in the illustration used for hydrochloric acid, Experiment 16. Collect the gas over water. *Heat gently. As soon as rapid action begins, remove the burner and allow the action to continue by itself.* Test the gas with the flame, and note its appearance. Test the bottles mouth up.

B

What are the characteristics of nitrogen?
Has any especially characteristic test for nitrogen been given here?
What is the especial use of nitrogen in the air?

Exercise 26. Study of Ammonia Gas. — Our ordinary household *ammonia* comes chiefly from the gas works, where one of the products of the distillation of coal yields *ammonia water*. Ammonia water is really ammonium hydroxide, and when this substance is heated *ammonia gas* passes off. If this gas is cooled or compressed sufficiently it forms *liquid ammonia*. Liquid ammonia, which is ammonia gas liquefied, is a compound entirely distinct from ammonia water, which may be formed by ammonia gas running into water and forming ammonium hydroxide. *Do not inhale the gas.*

Apparatus — Flask, two-hole stopper, thistle tube, and right bend, wide-mouth bottle, pneumatic trough.

Place about 25 cc. of strong ammonia solution in a flask. Fill the trough two-thirds full of water. Set up the apparatus in a similar fashion to that used in the experiment with hydrochloric acid (Experiment 16) *excepting* that instead of the tube leading downward we use a right bend of glass tube pointing upward. (Such a bend is shown at the extreme left of the illustration for preparing hydrogen.) Place the generator on a ring stand and heat *gently*, collecting the gas in a test tube by upward displacement. When the tube is filled with the alkaline gas, as shown by the change in color of a piece of moist red litmus held near the mouth of the bottles, close it and test for *solubility* in the pneumatic trough.

Hold a piece of litmus and a piece of turmeric loosely against the exit tube. Note results. Try to describe the *odor* of the gas. Fill the bottle with the gas and try the *flame test for a gas*.

Mix together a little powdered lime and powdered ammonium chloride in a test tube. Heat and hold litmus and turmeric above the tube. Note the odor of the gas. How is ammonia formed here?

B

What are the four chief characteristics of ammonia gas?

What is the most characteristic test for ammonia gas?

How would you know that there was ammonia in solution in a bottle of ammonia at home?

How does this gas compare in solubility with the other gases which you have studied?

You have now studied the following gases: oxygen, hydrogen, hydrochloric acid, chlorine, carbon dioxide, nitrogen, ammonia.

Fill out the following table from your notes, copy it in your notebook, and submit it before going farther.

(You may answer the questions where possible by "yes" or "no.")

	Burn?	Support Combustion?	Prevent Combustion?	Acid, Alkali, or Neutral?	Most Characteristic Test?
Oxygen					
Hydrogen					
Hydrochloric acid					
Chlorine.............					
Carbon dioxide.......					
Nitrogen					
Ammonia					

By reference to your table how could you tell whether an unknown gas in a bottle was oxygen, hydrogen, hydrochloric acid, chlorine, carbon dioxide, nitrogen or ammonia?

Exercise 27. Oxidizing Action of Nitric Acid. Formation of Sulphuric Acid from Nitric Acid. Test for Sulphuric Acid. — The experiment which follows is not given so much to illustrate a process for making sulphuric acid, although oxides of nitrogen are used to oxidize sulphur dioxide gas in the manufacture of sulphuric acid, as to illustrate the *oxidizing* power of nitric acid. The process of *oxidation* plays many important parts in our every-day life. Oxidation goes on, of course, every time gas or coal, wood or metal, burns in oxygen, but there are many other kinds of oxidation. Our common alcohol may be oxidized to the acetic acid of vinegar. Oxidation can take place without oxygen appearing at all. With the two iron chlorides, for example, ferrous chloride, $FeCl_2$, and ferric chloride, $FeCl_3$, when the ferrous form changes to the ferric form oxidation is said to take place. The iron atom in the ferric form is able to combine with one more atom of chlorine than it can enter into combination with in the ferrous form. Chlorine may be considered in this case an equivalent to oxygen, as two atoms of chlorine can replace one of oxygen.

To obtain sulphuric acid, sulphur dioxide must usually be oxidized to sulphur trioxide. When this is added to water, sulphuric acid is produced. $SO_3 + H_2O = H_2SO_4$. This is the action which is going on in this experiment. Simply note the change caused by the barium chloride test. We shall consider this test more carefully later on.

Use only a small pinch of sulphur and use the concentrated nitric acid. *Be careful in handling the concentrated acid.* Be sure that no sulphur gets into the sinks. *Do this experiment under a hood.*

Apparatus — Test tubes.

Place a small quantity of sulphur in a test tube and add to it a few cc. of concentrated nitric acid. Take the test tube to the hood and heat to boiling, holding the mouth of the tube away from you and bringing the acid *just to boiling.* As soon as the active boiling occurs *remove the tube from the burner,* but continue to hold it under the hood for a few moments until the gas evolved has nearly ceased to flow. When the flow of gas has practically ceased, bring the tube carefully back to your desk and place it in the test-tube rack.

Make a dilute solution of *sulphuric acid* by adding about one cc. of dilute sulphuric acid to a test tube full of water. (This is a test solution to show the effect of adding barium chloride solution to sulphuric acid.) Place the test tube containing this new solution beside the test tube containing the solution you have just boiled, and add a little barium chloride solution to both. Note the results and record.

B

What evidence is there that sulphuric acid was formed?
What chemical changes took place in oxidizing the sulphur?
Explain oxidation.

Exercise 28. Nitrates. Two Tests to Determine the Presence of Nitrates. — We shall see later that there are two kinds of analyses, qualitative and quantitative analysis. All the work we have been doing up to the present time has been qualitative testing, studying the properties of substances and their action when treated with certain reagents. Qualitative analysis differs from qualitative testing in that it separates out from a solution, or proves to be present in a mixture or compound, certain elements or radicals. Here two tests are shown to determine the presence of the radical of nitric acid (NO_3). Nitrates, compounds which contain this radical, are among the valuable fertilizers of the ground. Nitrates are regularly produced by the action of nitric acid on metals or metallic compounds.

Apparatus — Test tube, test tube holder.

Take two test tubes and place in each a couple of crystals of potassium nitrate. Dissolve in about 10 cc. of water. Add about 2 cc. of concentrated sulphuric acid to each tube. This will change the nitrate into free nitric acid. Drop into the first tube a couple of pieces of *copper filings.* Note results. Heat if no change occurs. Copper nitrate is formed.

To make sure that the conditions which appear in this test are not the result of the sulphuric acid which was added, pour 2 cc. of concentrated sulphuric acid into 10 cc. of water and add a few copper filings.

The second test must be performed carefully to make it a success. Pour *gently* a solution of ferrous sulphate on to the solution in the second test tube, so gently that the ferrous sulphate (iron sulphate) will float on top of the first solution. Note the action at the point where the two liquids meet in the tube.

B

What would you do to test an unknown substance for the presence of a nitrate by means of the copper test?

How do you know that this result was not due to the sulphuric acid?

How would you test an unknown substance for a nitrate by means of the ferrous sulphate test?

What salt is formed by the action of copper on nitric acid?

Exercise 29. Proportion of Oxygen and Nitrogen in Air. — In this experiment we are to consider nitrogen and oxygen with especial relation to the *proportional amount of the two gases* which exist in the air. The use of yellow phosphorus in this experiment makes *great care* essential if it is to be performed as a class exercise. It is often wise to perform it as a lecture-table experiment.

In this experiment we prepare nitrogen by a second method, burning the oxygen out of the mixed oxygen and nitrogen of a body of air enclosed in a bottle by means of yellow phosphorus, an element which oxidizes with extreme rapidity. If we burn out the oxygen of a bottle full of air by the use of yellow phosphorus there will be left in such a bottle nitrogen, which does not combine with phosphorus, and an oxide of phosphorus left after the burning. Of these two substances the oxide of phosphorus is soluble, the nitrogen is not. In consequence, as the action takes place over water, the oxide of phosphorus dissolves, water rises to take its place, and nitrogen is left as a gas. As the water occupies the volume of the phosphorus oxide, and the phosphorus oxide occupied the volume of the oxygen, the water which rises is equal in volume to the oxygen which was present at first. The volume of water, then, gives us the volume of oxygen that was present in the bottle. If we first measure the whole volume of the bottle and then measure the amount of water which has risen into the bottle, subtracting the volume of water from the whole volume will give us the volume of nitrogen. Divide the volume of nitrogen by the volume of oxygen and we shall have the *proportional volumes* of nitrogen and oxygen in air. We may neglect the presence of any of the other constituents of air here. Nitrogen and oxygen are the two parts of air which concern us most.

14. This illustration shows a bell jar, within which is phosphorus burning on a crucible cover. The experiment can be performed satisfactorily with a wide-mouth bottle.

Apparatus — Two wide-mouth bottles, crucible cover, forceps, pneumatic trough.

Part I. As a preliminary step to show that nitrogen is prepared by this method and that the oxide of phosphorus formed dissolves, burn a small piece of phosphorus placed on a crucible cover under a bottle of air, placed over water in the pneumatic trough. Test the gaseous product left after burning by the flame test.

Part II. Measure the capacity of a dry clean bottle in cubic centimeters, record, and repeat the experiment just performed. Leave the bottle full of gases in the pneumatic trough and endeavor to just seal the top with water. Allow the water to rise in the bottle as far as it will go. The rising water takes the oxide of phosphorus, formed by the burning, into solution and leaves the nitrogen free. When the water has ceased to rise, slip a glass plate under the bottle, reverse the bottle carefully, allowing *none of the water* in it to escape, and measure the number of cc. of water that have entered. This should show the amount of *oxygen present* that was converted to oxide. Subtracting this amount from the total capacity of the bottle shows the *volume of nitrogen* left.

B

How do the proportional parts of oxygen and nitrogen in air which you obtained correspond to the real proportions in which they are found in the atmosphere?

NOTE. — Be sure that at the end of this experiment every piece of phosphorus in the laboratory is safely in the phosphorus bottle. Take no chances with phosphorus.

Exercise 30. Water Vapor in the Air. Efflorescence and Deliquescence. Water of Crystallization. — Water of crystallization entering into the composition of a dry solid generally changes it to a crystal of definite shape. This *water of crystallization* becomes a part of the molecule, so that while we write the formula of dry or anhydrous sodium sulphate as Na_2SO_4, we write the formula of the crystallized salt as $Na_2SO_4 \cdot 10\ H_2O$. The latter expression indicates that ten molecules of water have become a part of the molecule. The *water vapor* of the atmosphere is one of the most necessary of the compounds and elements present in that mixture. It furnishes necessary protection to man and beast against the burning rays of the sun, acts as the medium which holds and carries the rain, and in sun and shadow gives moisture to plant life. Many chemicals have the power to either lose water to the air around or to take it from the surrounding atmosphere.

An *efflorescent* body is one which, on contact with the air, loses water which it contains. A *deliquescent* body is one which has the property of absorbing water vapor from the air.

15. The Mercury Barometer. — By means of this instrument we read the varying pressure of the atmosphere. By the use of Boyle's Law and the knowledge given us by the barometer, we may reduce any volume of gas measured under any conditions of atmospheric pressure to standard conditions of pressure.

Apparatus — Evaporating dish, asbestos, two watch glasses.

Part I. Take a crystal of copper sulphate, place it in an evaporating dish and heat it carefully until the color has wholly disappeared. Let it cool and drop a few drops of water on it.

B

On what does the color and shape seem to depend?

Can you give any explanation of this action?

Remember the result in order to compare it with future experiments.

Part II. Leave a piece of calcium chloride on a watch glass under your desk for two or three days, examining it each time you come in. Leave a piece of crystallized sodium carbonate on a watch glass for two or three days, examining now and then. Note the result. Explain the action.

B

Where does the water come from?

We often need to remove water vapor from gases we are collecting; how may we dry them effectually?

Exercise 31. Boyle's Law. — In this experiment we are to consider the law which governs the relation of pressure to volume in gases measured at a constant temperature. Gas volumes vary both with changes of temperature and pressure. In consequence, if we wish to consider the effect of changes of pressure alone upon the volume of a gas we must keep the temperature *fixed* or *constant.* This is done readily enough by taking the measurements at the temperature of the room, which remains sufficiently constant for all practical purposes during the progress of the experiment. The somewhat crude methods of measurement used with the ordinary Boyle's Law tube will not give exactly equal results in the final ratio, but a little care will show the relation between pressure and volume clearly.

Set up the Boyle's Law tube in the fashion shown in the tube at the left of illustration 15. The column of air in the closed end of the tube is acted upon by a pressure equal to the height of the mercury in the long arm of the tube above the level of the mercury in the short arm. That is, the actual pressure of the mercury is equal to the difference in level in the mercury in the two arms. There are two factors to be considered when the temperature is constant as here, pressure and volume. To obtain the total pressure we must add the pressure of the atmosphere, which is constantly bearing down upon the surface of the mercury in the long arm, to each individual pressure of mercury. We obtain the *atmospheric pressure* by reading the *barometer.* We obtain the *mercury pressure* by reading the *difference in level* between the mercury in the two arms. We obtain the *volume of air* in the tube by measuring the *length* of the column of air. This is the only record we need to take for volume, as the breadth and thickness of the tube are constant and the volume varies as the length. All measurements are to be taken in centimeters and tenths of centimeters. A model set of measurements would read as follows:

Atmospheric Pressure (Barometer Reading)	Mercury Pressure (Difference in Level of Mercury)	Total Pressure	Volume.
76.9 cm.	13.5 cm.	90.4 cm.	22 cm.

This experiment is often best treated as a lecture-table experiment. The experiment on the law of Charles, while simpler, is also a satisfactory lecture-table experiment.

Apparatus — Boyle's Law tube.

Adjust the apparatus and pour in sufficient mercury to have an equal level in each of the arms. Read and record the length of the column of air. Now add carefully through a funnel four successive portions of mercury, adding approximately 15 to 20 cm. in height each time. Read and record the volume of air after each addition. Read and record the difference in level of the mercury each time you add the liquid. Arrange your data in the following form:

Pressure = P	Volume = V	P × V
I. Atmospheric pressure.	Volume I. =	Pressure I. × Volume I. =
II. Atmospheric pressure + pressure A. =	" II. =	" II. × " II. =
III. Atmospheric pressure + pressure B. =	" III. =	" III. × " III. =
IV. Atmospheric pressure + pressure C. =	" IV. =	" IV. × " IV. =
V. Atmospheric pressure + pressure D. =	" V. =	" V × " V. =

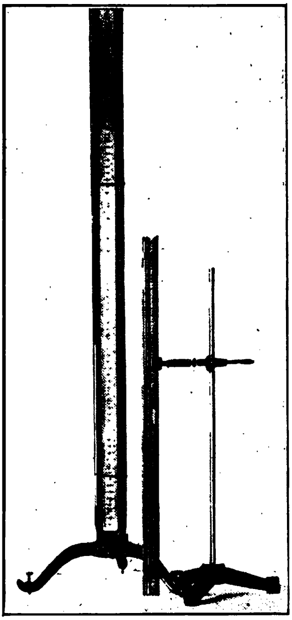
Note that pressures, A, B, C, D, are the additional pressures of mercury obtained by pouring mercury into the long arm of the tube and that they equal the respective differences of level in the two tubes. When this work is completed multiply each pressure by its corresponding volume and enter in the column marked P × V. Compare the results obtained. Now, according to the laws of proportion, if we have four quantities a, b, c, and d, and $ad = bc$, $a:b = c:d$. That is, if the product of two numbers, a and d equals the product of two other numbers, b and c, either two may be made the means of a proportion in which the other two are made the extremes. Considering this in terms of pressure and volume we shall find a valuable law. If the experiment has been performed correctly, the product of each individual pressure times volume should equal the product of every other pressure times the corresponding volume. Pressure × Volume under such circumstances becomes a constant factor, and we may say, using P for first pressure, P′ for second pressure, V for first volume and V′ for second volume, $P \times V = P' \times V'$, Then, by the rules of proportion just given, $P : P' = V' : V$. That is, the first pressure : second pressure : : second volume : first volume.

B

Work out the proportions obtained in this experiment and see how nearly they approach the law.

Exercise 32. Law of Charles. — We know by the Law of Charles that all true gases under the same conditions of pressure expand by the same amount for every degree that the temperature rises, or, to put it in another way, that the same expansion of two like volumes of gases would result when these like volumes were heated and rose the same number of degrees of temperature. We are to determine here what the amount of this expansion is for the specific case of air heated at constant pressure. Heating the tube of air in one position, an upright one, keeps the pressure constant.

The illustration for the two experiments shows on the right a glass tube holding a column of dry air held in by a bit of mercury and mounted on a section of a meter rod. This is one form of tube used for this experiment. There are other forms which are satisfactory. To obtain the lower temperature the column of air may be measured at the temperature of the room, or, if desired, the tube may be packed in ice and a temperature of approximately 0° C. obtained. To obtain the higher temperature the column of air should be measured in steam. This gives approximately 100° C. The tube can be heated most easily in the thermometer tube of a Gilley boiler or of the heating apparatus commonly listed as "Apparatus A."



16. APPARATUS FOR BOYLE'S LAW AND CHARLES' LAW. — On the left is one type of Boyle's Law Tube, consisting of a short and long arm. The short arm is closed and contains a volume of air, the long arm contains the changing column of mercury. For convenience in reading the tube is mounted on a meter rod. At the right is shown a form of Charles' Law tube described in Exercise 32.

Apparatus — Charles' Law tube.

Measure the length of air column when held perpendicularly. Do not heat the tube by touching it with your hand. Note the temperature of the room. This will be the temperature of the column of air. Place the tube in a boiler and expose it to steam. When steam is coming off rapidly, measure to a tenth of a centimeter the length of the air column. Take your measurements from the bottom of the mercury to the inside of the top of the closed tube and record. Leave the air tube in the boiler until the length of the column is constant, as shown by the fact that with three readings, taken at intervals of two minutes each, the length remains unchanged. Take the temperature of the steam and record as follows:

1. Original length of air column =
2. Length of air column after heating =
3. Expansion of air column affected by heating =
4. Temperature after heating =
5. Temperature before heating =
6. Rise of temperature =
7. Amount of expansion for one degree Centigrade =

To obtain this final result, as a preliminary step, divide the expansion by the original length of air column. Remember the lengths of the column are the only measurement of the column taken. This is because the diameter of the cylinder of air in the tube is always the same and the only factor affected by heat (as was the case with pressure) is the length. Putting it still another way, volume varies directly as length. Therefore we measure length to give us results concerning volume. Your final result gives you the percentage of expansion, as compared to the whole original volume. Divide this result by the number of degrees of temperature through which the air rose. This gives you the percentage of expansion of the whole volume during a rise of one degree Centigrade in temperature.

B

The percentage of gas expansion which you have obtained experimentally here, should correspond reasonably closely to the actual percentage by which any true gas will expand for a use of one degree Centigrade in temperature. That is, it should expand $\frac{1}{273}$ of its original volume for each such rise of 1°. Change your decimal result to common fractions and compare it with the true percentage. Perform this operation as in the model example which follows.

Change .00366 to common fractions:

$$.00366 = \frac{366}{100{,}000} \div 366 = \frac{366 \div 366}{100{,}000 \div 366} = \frac{1}{273}$$

Exercise 33. Silicic Acid. — Silica (silicon dioxide), the main constituent of sand, is especially valuable for making glass. The ordinary glass we use in houses is made from fine clean sand combined with other compounds. In this experiment we take silicic acid, derived from the element silicon, from a compound known as water glass, and then reduce this silicic acid (by heating) to silicon dioxide and water with the aid of strong hydrochloric acid. Great care should be taken not to spill the acid and it would be well to do this experiment *under a hood.*

Apparatus — Evaporating dish.

Add sufficient strong hydrochloric acid to a small quantity of water glass to produce an acid reaction. Note results. After action has ceased, evaporate the substance left in the dish to dryness. The jelly-like mass formed is silicic acid. Salt is formed at the same time. The reactions may be written as follows:

Sodium silicate (water glass) + hydrochloric acid = silicic acid + sodium chloride (salt).

Silicic acid (heated) = water + silicon dioxide.

The white powder left is a mixture of silicon dioxide and salt. Note that the precipitation of silicic acid in this experiment depends upon the use of concentrated solutions of water glass and hydrochloric acid. Dilute solutions do not precipitate. In the use of concentrated acid *great care* should be taken not to spill a drop.

B

Discuss the use of silicon in the soil and in glass-making.

Exercise 34. A Test for Silicon in its Compounds. — In this experiment we are to test for the silicon which exists in quartz. The quartz used here is like the ordinary quartz of pasture rocks.

Apparatus — Ignition tube, platinum wire.

Mix equal parts of powdered quartz and powdered fluorspar, calcium fluoride, with a drop of concentrated sulphuric acid in an ignition tube. Heat, holding a drop of water in a tiny loop of platinum wire within the mouth of the tube. Observe the action on the water. The mass formed in the loop is produced by a gas, a compound of silicon and fluorine acting upon the water, and is called a hydrated compound of silicon.

B

Here is a test for silicon compounds. Briefly review the different types of tests that we have tried for solids, liquids, and gases.

Exercise 35. Study of Sulphur. — Your commonest knowledge of sulphur comes probably from the sulphur match. The ease with which sulphur melts is shown every time the yellow solid on the side of such a match turns to brown liquid. The ease with which sulphur oxidizes is shown when it burns. It is one of the oxides of sulphur, sulphur dioxide, which gives a sulphur match its characteristic smell.

In this experiment, as in all the others, try to use as many of your five senses as possible in determining what the properties of the element sulphur may be. Note that sulphur unites with iron, forming iron sulphide in a more or less similar fashion to the way in which oxygen unites with iron, forming iron oxide.

Apparatus — Test tube, beaker.

Heat a small amount of sulphur in a test tube, preferably using a tube which has been used before for the same purpose. Note the ease or difficulty with which it melts. Note also the varying consistency of the liquid in the tube as the heating continues. After the sulphur has boiled note the appearance of the solid formed at the top of the test tube.

Then take a piece of iron wire about one foot long from a spool of wire, and heat one end red hot in the flame. Boil sulphur at the same time in the same flame. Plunge the red hot iron into the sulphur vapor and note the effect of the combination.

B

What are the properties of sulphur as regards comparative weight with water, color, ease or difficulty with which it changes its state?

What are the changes in the liquid sulphur?

What compound is formed when the hot iron wire is plunged into the sulphur vapor?

Exercise 36. Preparation of Plastic Sulphur. — In this experiment we are to form one of the allotropes of sulphur, plastic sulphur. Look up the definition of an allotrope if you are not sure of it before commencing the experiment.

Apparatus — Funnel, test tube, pneumatic trough.

In preparation for this experiment fill your pneumatic trough nearly full of water and place your funnel upside down in the water so that it stands on a firm base. To obtain the plastic sulphur we are to pour liquid sulphur into water. It is easier in many ways to pour it on to the glass of a funnel placed under water.

Heat a few grams of flowers of sulphur in a test tube until the element passes through the thick liquid state in which heated sulphur is found and passes from there into the thin, easily poured state. When you have heated it to this point it should be boiling. When the sulphur is boiling, *carefully* pour the sulphur in a slight stream round and round the glass of the funnel under the water. After cooling note the product obtained. When it has wholly cooled take it up in your hands.

B

How does this form of sulphur compare with the other allotropes of sulphur?

Note. — Bring a flower to school for the next experiment. Choose any color except white.

Exercise 37. Study of Sulphur Dioxide Gas by Burning Sulphur in Air. — Sulphur dioxide gas is a comparatively mild bleacher of especial value for the bleaching of straws, silks, and other substances which would be injured by chlorine. Test the gas with a view to determining all the characteristics you can possibly find.

Apparatus — Pneumatic trough, test tube.

1. Burn sulphur in a deflagrating spoon held in a bottle of air which has been partly closed by means of a watch glass.

Note the appearance of the product. Note its effect on wet blue litmus paper. Wet a flower and hold it for some time in the bottle full of gas. Note the change.

2. Fill a second bottle and place in it 15 cc. of water. Close and shake well. Note any change in appearance. Test with litmus.

3. Burn sulphur in a deflagrating spoon until the fumes come off rapidly, then plunge the spoon into the bottle and hold it there until the bottle seems full of gas. Quickly place the bottle full of gas mouth downward in a pneumatic trough half full of water. See if any changes in the water level occur. This is a common test for the *solubility* of a gas.

B

What elements must be in the product from the sulphur?

Name this product in two ways.

What evidence exists that the gas is soluble in water?

What did the litmus paper test show?

Exercise 38. Study of Hydrogen Sulphide Gas. — This is the gas which, produced as eggs grow stale, gives the characteristic odor of rotten eggs. *Hydrogen sulphide* is also produced by many varieties of burning coal and of coal gas. This is the gas which commonly blackens the white lead paint of houses and tarnishes silver and gold ornaments.

Hydrogen sulphide gas is of especial value in qualitative analysis for the separation and detection of various metals such as mercury, antimony, lead, tin, bismuth, arsenic, copper, and cadmium by the formation of their sulphides. In the separation of different metals in qualitative analysis hydrogen sulphide gas is bubbled through solutions which contain compounds of these metals. In this experiment we are to consider certain properties of the gas and also its action on a compound of one of the metals, lead. It is most essential that the gas should be kept continuously *under the hood.*

Apparatus — Gas bottle, two-hole stopper, thistle and delivery tubes.

Do this under the hood and do not inhale the gas.

Place a small piece of iron sulphide in a gas bottle and set up a generator under the hood. Fill a beaker full of water, and connect with the generator by means of a delivery tube. Keep the gas running into this when not otherwise using it, thereby forming a solution of the gas. Save the solution.

Dissolve a few crystals of lead nitrate in a test tube full of water and place it with the rest of the apparatus in the hood. Then, when everything is ready, and *not before*, pour in sufficient hydrochloric acid through the thistle tube to cover the iron sulphide and the bottom of the thistle tube. Collect two test tubes of *hydrogen sulphide gas* by downward displacement, and test one with the *flame* test. Test for *solubility* by placing the other tube mouth downward in the pneumatic trough as you did in the previous experiment with sulphur dioxide.

Run the gas into the test tube containing the *lead* nitrate for a minute and note results. When the period is over, open the windows, close the door and take the generator rapidly to a sink. *Do not inhale the gas.* Wash thoroughly and return any iron sulphide remaining, to the bottle marked "Iron Sulphide Residues." Preserve the solution in the beaker. Finally take a little more lead nitrate solution and test it by adding the solution of hydrogen sulphide already formed. Note results.

B

What are the chief characteristics of hydrogen sulphide gas?

Write the reaction for forming the gas.

What was formed by the action of the gas on lead nitrate?

Write the equation for the action of hydrogen sulphide on lead nitrate?

How about the solubility of the gas?

Could you see any difference between the action of the gas itself and of its solution in water?

Exercise 39. Study of Sulphurous Acid and one of its Salts, Sodium Sulphite. — The salts of sulphurous acid are chiefly of value commercially for strengthening the fibers of wood, so as to make them available for paper-making. For obtaining sulphurous acid in quantity the preparation from copper filings and sulphuric acid given here is a satisfactory laboratory method. It is better to do this experiment *under the hood.*

Apparatus — Flask, delivery tube, thistle tube, bottle, litmus.

Place a small quantity of copper filings (sufficient to just cover the bottom) in a flask. Fit up the flask with thistle and delivery tubes, and pour through the thistle tube enough concentrated sulphuric acid to cover the copper completely. (Be very careful in using the *concentrated* acid not to *spill a drop.*) Set up the apparatus in the same fashion as shown in the illustration in Ex. 16 on hydrochloric acid.

Heat carefully, and form *sulphurous acid.* Pass the gas into a bottle full of water. Test with litmus. When the water will take up no more gas, stop heating.

Filter the residue in the flask, noting its appearance, and throw the remaining solid into the waste jar.

Neutralize a small quantity of sulphurous acid in an evaporating dish with sodium hydroxide, and evaporate a couple of cc. to dryness in a test tube. If time permits repeat using potassium hydroxide.

B

What is the difference between sulphurous and sulphuric acid?

What is the difference between a sulphate and a sulphite?

Write the reaction which takes place between the sodium and potassium hydroxides and the sulphurous acid.

Reaction of H_2SO_4 on copper was as follows: $Cu + 2H_2SO_4 - CuSO_4 + 2H_2O + SO_2$.

Exercise 40. Study of Sulphuric Acid. — In this experiment we are to consider three things: the hygroscopic action of sulphuric acid; the heat evolved when sulphuric acid is added to water; the barium chloride test for sulphurous and sulphuric acids, for sulphites and sulphates. A hygroscopic substance is one that has the power to remove water from other substances containing it with which it comes in contact. Many of the commercial uses of sulphuric acid such as its use in the preparation of glucose and of sulphuric ether depend upon the hygroscopic action of the acid. Sugar contains carbon, combined with hydrogen and oxygen. These elements are present in the proportions in which they are found in water. We are to consider the action of the acid on sugar from this standpoint of hygroscopic action. Later on we shall consider it from a different point of view.

The heat absorbed and evolved in chemical reactions is of such importance that a whole branch of chemistry, thermo-chemistry, is devoted to this subject. We have already seen an example of heat absorption in the mixture of ice and salt, which was like the ice and salt in the ice cream freezer. Here we take up an example of heat evolution.

We have already tried the barium chloride test for sulphuric acid, when we oxidized sulphur by the use of nitric acid. We are to use this reagent here for sulphurous and sulphuric acids and for their salts to show the difference in reaction between the forms.

Concentrated sulphuric acid is used here. It is the oil of vitriol of commerce. Properly handled it is entirely safe. Remember that carelessly handled it is very *dangerous.*

Apparatus — Test tubes, thermometer.

Cover the bottom of a test tube with sugar and pour on to it sufficient concentrated sulphuric acid to cover it. Let it stand for a while and interpret the results.

Part II. Fill a test tube one-third full of water. Take a thermometer, read the temperature of the water and record. Add an equal amount of concentrated sulphuric acid to the water in the tube, taking care not to spill a drop. Take the temperature again and record. Note the change.

Part III. Make solutions of dilute sulphuric and sulphurous acids in two test tubes. Add a little barium chloride solution to each, noting carefully which tube contains sulphurous acid and which contains sulphuric acid. Now add hydrochloric acid. If no change occurs add a little more concentrated hydrochloric acid to each. Note changes and record. Repeat this test in all its details with dilute solutions of sodium and potassium sulphites and sulphates. Record the results obtained from the four salts and two acids in a table. Notice that in these cases we are comparing sulphites and sulphates

B

What is meant by hygroscopic action? Have you met any similar action to this before?

What did the sulphuric acid do to the sugar?

What was the temperature change? Have you met any similar change in daily life or in the laboratory?

How can you differentiate between sulphurous and sulphuric acids? between sulphites and sulphates?

Exercise 41. Study of Yellow Phosphorus. — The following experiment with yellow phosphorus requires great care in the hands of laboratory sections on account of the fire risks. The distribution of small pieces of yellow phosphorus about the ordinary laboratory with its wooden floor and wooden desks entails especial caution.

Some knowledge of yellow phosphorus, however, is most advisable. This is a satisfactory lecture table experiment.

In this experiment we are to consider a few of the characteristic properties of yellow *phosphorus*, one of the two allotropes of this element. This is the phosphorus of the *common sulphur match.* From this element are derived the *phosphates* so necessary for plant life.

Apparatus — Two evaporating dishes, beaker, iron wire, iron plate, forceps.

Hold a small piece of *phosphorus* in forceps in air. Observe the difference between the color of the inner part and the outer coating. This outer coating is an oxide. The smoke of the oxide of phosphorus formed shows rapid oxidation.

Take two pieces of phosphorus about the size of a pea in two separate evaporating dishes. Cover one with lampblack, leaving the other piece exposed to the air. Note the action of the lampblack on phosphorus, due to the coating which changes the kindling point. With a heated wire set fire to a small piece of uncoated phosphorus and place a beaker over the thick white smoke which rises. After the action is over, fill the beaker with water and test with litmus paper.

Place a small piece of phosphorus about the size of a pea in an evaporating dish and drop on it about an equal amount of iodine crystals. Note the results. Place an iron plate on a ring stand and place small bits of red and yellow phosphorus on opposite sides of the plate. Heat the plate gently, and continue heating until both the red and yellow phosphorus have taken fire. Compare the relative inflammability of the two.

B

How does the inflammability of phosphorus compare with that of the other substances you have studied?

What is the special value of phosphorus in a match?

What is formed when phosphorus burns in air?

What did the test with litmus show?

When sulphur is burned in air and added to water, what is formed?

When phosphorus is burned and added to water, what is formed?

What does the union of phosphorus with the oxygen of the air and with iodine indicate as regards the ease with which this element enters into combination?

Compare the relative inflammability of yellow phosphorus and red phosphorus.

Exercise 42. Study of Red Phosphorus. The use of Red Phosphorus in Safety Matches. — In the safety match, *red phosphorus* may be on the match, while some *oxidizing agent*, rich in oxygen, such as *potassium chlorate*, may be placed on the box. The union of the two produces the flame. In this experiment we are to change *red phosphorus* to *yellow phosphorus* and also to observe the difference between the action of red phosphorus, used by itself, and in union with some oxidizing agent, some substance that will yield oxygen. We are also to form a sublimate.

Take an ordinary piece of glass tubing about a foot long and close it at one end. After the glass has cooled, place a small quantity of red phosphorus in the tube by folding a filter paper, piercing its center with a pencil point and then pouring a little of the red phosphorus powder into the tube through the hole. Heat continuously in the flame and note the *sublimate* which forms. After a considerable amount of sublimate has been formed on the upper part of the tube, *and the tube has wholly cooled* scratch the coating of sublimate with an iron wire. *Take great care here. Do not touch the coating with your fingers. Drop the tube into a separate waste jar when you finish. Yellow phosphorus is formed.*

Part II. Take an iron plate and repeat one of the processes of the last experiment, 41, with this difference. Instead of comparing the two types of phosphorus we will consider the action, under the influence of heat, of red phosphorus alone and with an oxidizing agent. Place on one side of the plate a small quantity of red phosphorus and on the other an equal quantity of the same substance mixed thoroughly with about the same amount of potassium chlorate. These two specimens will represent the different conditions which exist when red phosphorus is used alone and when it is used with an oxidizing agent like potassium chlorate. Heat gently as before and note results.

B

What does the sublimate formed from red phosphorus show?
What is the principle of the safety match?
Why are safety matches better than the common matches?

Exercise 43. Effect of the Electric Current upon Solutions that are Electrolytes. — Tableware that is electroplated has its exterior metallic coating, which is generally silver, deposited upon it by an action similar to that shown in Part I. The action in Part II illustrates the ionic conditions of a conducting solution during the passage of a current. The illustration placed opposite this paragraph shows a single fluid galvanic cell generating a current.

17. Generating Electricity. — A single fluid galvanic cell. The current passes from the zinc on the left to the copper on the right through sulphuric acid, which is broken down by electrolysis. The current then passes from the copper through the wires and back to the zinc outside the cell.

In the illustration placed between this experiment and the next is shown a single fluid galvanic cell representing a generator of electricity. In practice the electric current furnished a laboratory section usually comes from a battery made up of several cells or from a dynamo or motor generator. A single fluid cell is used here simply as a type of generator. In any case the electrodes pass in and out of the solution to be tested in the manner shown in the wires passing into and out of the beaker. A double connector connecting two separate wires is shown in the foreground at the right. The instrument at the right, a galvanoscope, is not used in this experiment. Copper electrodes are used in the first part of the experiment. Platinum electrodes are used in the second part. For discussion of the passage of an electric current through a solution consult the text-book.

In all these experiments take care not to touch the electrodes to each other, thereby short-circuiting the current. Do not let the binding posts drop into the solution.

Apparatus — Beaker, connectors, two pieces copper wire, two double connectors.

PART I. EFFECT OF AN ELECTRIC CURRENT ON A SOLUTION OF COPPER SULPHATE

Pass the current through a solution of copper sulphate held in a beaker. Note results. After allowing the current to run a short time, reverse the current by disconnecting the wires leading from the generator and exchanging them so that the positive wire from the battery is now connected with what has been the negative electrode.

B

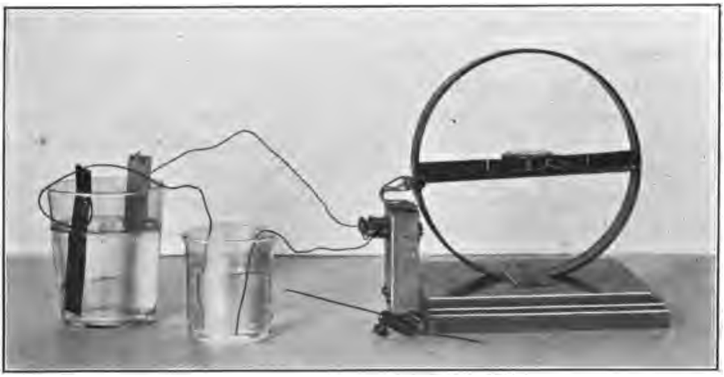
What effect is produced when the current acts on the copper sulphate?

What happens when the current is reversed?

Why does this happen?

PART II. EFFECT OF AN ELECTRIC CURRENT ON A SOLUTION OF SODIUM SULPHATE.

Same apparatus as in Part I, except that you are to use platinum electrodes. Return these electrodes *without fail* after use.



18. APPARATUS DESCRIBED IN EXERCISES 43 AND 44.

Pass the current through a solution of sodium sulphate containing phenolphthalein. This solution is made by adding a little sodium sulphate to water containing five drops of phenolphthalein. Note the changes which take place in the solution. You may disregard any metallic deposit which forms on the platinum.

B

What is formed at the positive electrode?
What is formed at the negative electrode?
Which is the anion and which the cation in this case?

Exercise 44. Comparison of Effect Produced by Passing Current through Solutions of Non-electrolytes and Electrolytes. — The galvanoscope, the instrument on the right in the illustration, shows the passage of an electric current. Connect the current directly to the outer binding posts of the galvanoscope and watch the swing of the needle. The principle of the galvanoscope is a matter to be considered in the study of physics. It is sufficient here to recognize that it shows the passage of an electric current by the swing of its compass needle. The needle ceases to swing and becomes stationary when the current ceases to flow.

Apparatus — Beaker, connector. Use platinum electrodes.

Connect the current through the galvanoscope and a beaker as shown in the illustration. Make a sugar solution by adding a little sugar to a beaker two-thirds full of water. Make a salt solution by adding a similar amount of salt to a similar amount of water. The sugar solution is a typical solution of a *non-electrolyte.* The salt solution is a typical solution of an *electrolyte.*

Pass the current into the sugar solution, connecting with the galvanoscope as indicated in the illustration, taking great care not to have the terminals which are in the solution, touch each other. Pass the current through the salt solution in a similar way. Note the action of the galvanoscope in each case. Then, remembering that the galvanoscope needle swings when a current passes into the wires, and remains stationary when no current passes, write down the results.

B

Explain, as best you can, what happens when an electric current passes through the solution of an electrolyte.

Explain the action when the current passes into the solution of a non-electrolyte.

Exercise 45. The Presence of Carbon in Fuel. — In this experiment we are to burn a candle with incomplete combustion, that is, we shall supply it with sufficient oxygen to keep a flame alive but not a sufficient quantity to produce complete combustion or burning. Such a condition exists when a lamp smokes. The smoking of a kerosene lamp is due to incomplete combustion, a condition which arises from an insufficient supply of oxygen. The oxygen in such a flame comes from the base holes of the lamp and must be provided to bring about the oxidation or burning of the compounds present in the kerosene. The carbon or soot produced by such incomplete burning is called lampblack.

Apparatus — Candle, wide-mouth bottle, crucible cover.

Light a candle under a bottle held in such a fashion that only a limited amount of air can reach the flame. Then burn it directly under a porcelain crucible cover, holding the cover with crucible tongs directly above the flame. Note results.

B

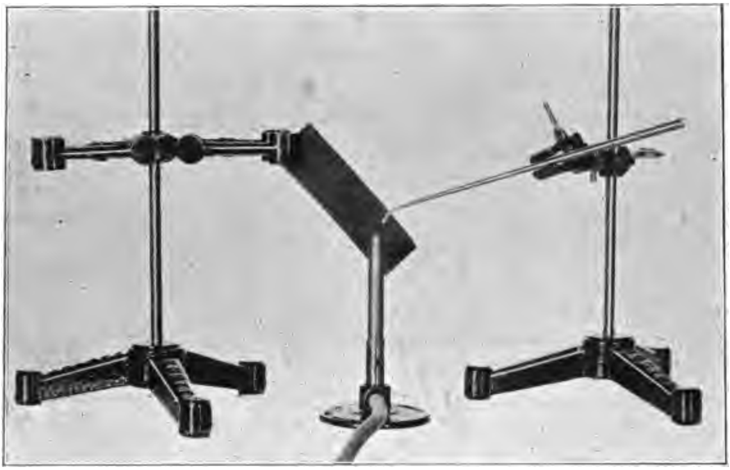
What elements seem to be present in the candle in this experiment?
Why does a lamp smoke?
What uses exist for lampblack?

Exercise 46. Reducing Action of Carbon. — In this experiment we are to consider a process largely employed for the reduction of metals from their ores. Such metals as iron and copper often appear in their mineral form as oxides. The action upon such metals of carbon, acting as a reducing agent, is shown here. This experiment can be performed in two ways. The method given here is perhaps the best, but if a powdered metallic oxide is mixed with powdered charcoal in an ignition tube, and heated thoroughly, the same result is obtained.

Reduction is a process opposed to oxidation. We reduce a substance when we remove from it oxygen or the equivalent of oxygen. In the following case we reduce lead oxide to metallic lead by removing oxygen and leaving lead. In similar fashion in our study of oxygen we reduced mercury oxide to mercury by driving off the oxygen. Reduction is possible without removing oxygen. If we remove the chlorine from iron chloride and leave metallic iron we shall still be reducing, for we remove chlorine, two of whose atoms are equivalent to one of oxygen.

Apparatus — Blow-pipe, block of charcoal.

Take a knife or a file and make a small hollow in one end of a block of charcoal. Into this hollow pour a small quantity of red lead (lead oxide, minium). Hold the block and blow-pipe in the fashion indicated in the illustration. Direct a steady blow-pipe flame upon the oxide, using the mouth as bellows and breathing through the nose. Continue until metallic lead forms. The metallic lead has been reduced from the oxide by the carbon during the heating. The coating which covers the lead as it cools is another oxide of lead, formed by the union of the metal with the oxygen of the air.



19. Use of the Blow-Pipe. — The relative position of the charcoal block, the burner, and the blow-pipe are shown in the illustration. The student, holding the block in one hand, directs the flame on to the hollow of the block by means of the blow-pipe held in his other hand.

B

What is the general use of carbon in removing combined oxygen from metallic ores?

What oxides of carbon exist? Under what different conditions do they form?

Exercise 47. Study of Marsh Gas. Methane. — In this experiment we are to form the simplest organic gas which exists, *methane*, the marsh gas of the autumn pool, a constituent of our ordinary coal gas, the fire damp of the mine. In organic chemistry methane assumes great importance as the first gas of the great series of the *hydrocarbons*. From this series of hydrocarbons are derived thousands of important organic substances.

The calcium oxide used in the preparation of methane is present chiefly for the purpose of drying out water which may be present in the sodium acetate or in the sodium hydroxide used. Sodium acetate contains two atoms of carbon and three hydrogen atoms beside sodium and oxygen. One of the carbon atoms and the three hydrogen atoms unite with the single hydrogen atom present in sodium hydroxide to form a compound made up of one atom of carbon and four of hydrogen, methane, CH_4.

Apparatus — Ignition tube, test tube.

Very thoroughly mix together 1 gm. sodium acetate, 2 gm. sodium hydroxide, and 4 gm. calcium oxide. Place enough of the powder in the ignition tube to about half fill it, allowing the powder to rest only on the lower side of the tube, leaving the upper half of the tube free.

Heat the tube and collect a test tube full of the gas over water. Test with the flame by carefully bringing a lighted match to the mouth of the test tube.

B

What are the characteristic properties of methane?

Where is methane found?

Exercise 48. Composition of Several Organic Substances. — The different gases which are formed in the use of our ordinary fuels by burning coal in stove or furnace, or by the burning of a match or gas flame, must go into the air which we breathe. It is a matter of no slight interest to us, therefore, to know something of what these products are. In this experiment we are to consider the products which are formed by the burning of several fuels. Since these fuels are all organic substances this experiment should also give considerable light on the problem of the composition of organic substances in general.

Apparatus — Two bottles, gas bottle, flask, watch glass.

Burn candle, match, a few drops of alcohol and a few drops of kerosene on watch glasses placed on asbestos, under cool, clean, dry bottles. Light the kerosene by means of your lighted burner. If the flame goes out in any of these cases it is because you have not given it sufficient air. Under such circumstances light the fuel again and hold the bottle higher. If you can use only one watch glass and one bottle, wipe both dry between each test. Examine the sides of the bottle carefully and record any change in appearance. Test all the gases formed with lime water (calcium hydroxide). Repeat this experiment with a bit of charcoal, brought to glowing heat in your burner flame and held in a bottle of air with crucible tongs.

B

What product of this combustion appeared on the sides of the bottles?
What product of this combustion did the lime water test show?
What elements are present?
What gases are generally produced by the burning of ordinary fuels?

Exercise 49. A Test for Alcohol by the Preparation of a Product of Methane, Iodoform. — In this experiment we are endeavoring to illustrate two things: first, to show a test for our *common alcohol* (spirits of wine), and second, to show the compound, *iodoform*, finally produced in this test, an important medicinal substance, closely allied to chloroform, and derived from methane. The production of this compound does more, therefore, than show the presence of alcohol. It illustrates the production of a derivative of methane. These derivatives are generally formed by the replacing of one or more of the hydrogen atoms of the molecule from which they are derived, by an atom or atoms of some other element. Halogen derivatives can be formed from methane by replacing hydrogen with chlorine, iodine, or bromine. In this case iodine is the replacing element. In iodoform three of the four hydrogen atoms of the methane molecule have been replaced by iodine atoms, forming iodoform, CHI_3, instead of methane, CH_4.

Apparatus — Beaker, thermometer.

Dissolve 4 gms. crystallized sodium carbonate in 25 gms. water in a beaker. Pour 2.5 gms. alcohol into the solution and, after heating to

80° C., as shown by the thermometer, gradually add 2.5 gms. iodine. Iodoform separates from the solution and may be recognized by its appearance and smell. Iodoform forms in yellow six-sided crystals and has a characteristic, clean, medicinal odor.

B

How could you test for alcohol in solution?
Give some of the uses of alcohol and iodoform.

Exercise 50. Study of Sulphuric Ether. — Sulphuric ether or ethyl ether is probably the most important anæsthetic that we possess. Its name is derived from the fact that it is produced by the action of sulphuric acid on alcohol. In the case of ether, as in the case of numerous organic compounds, the sense of smell gives us the simplest test by which we can recognize the substance. Ether is an excellent solvent, but the utmost care must be used in handling it on account of its great inflammability. *Be careful. This substance is very inflammable. Have no flame about. Before lighting a match be sure that the ether bottle is stoppered. Concentrated acid* is to be used in this experiment and the next. *Take care.*

Apparatus — Test tube, medicine dropper.

Part I. Preparation of Ethyl Ether (Sulphuric Ether). — Measure out into a test tube, with a medicine dropper, 10 drops of alcohol and 10 drops concentrated sulphuric acid. *Ether* is formed slowly. Let the liquid stand until you can note the peculiar smell.

Part II. Inflammability and Mobility of Ether. — Drop 6 drops of ether into a beaker. Cover closely with a watch glass and allow to stand for five minutes. Then cautiously bring a lighted match to the mouth of the beaker. Pour a few drops of ether into the palm of your hand, hold them there for a minute, and note result.

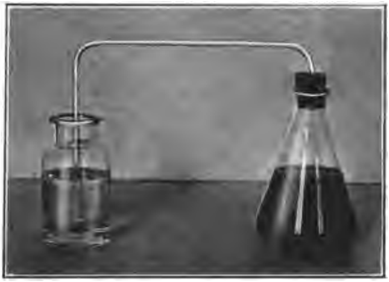
B

Define an anæsthetic.
What are some of the uses of anæsthetics?
What are the properties of sulphuric ether?

Exercise 51. Fermentation of Molasses to Impure Alcohol and Acetic Acid. — This experiment illustrates the action of yeast in producing a chemical change. A similar action takes place when sweet cider is changed to hard cider. The sugars of molasses, like the sugars of pressed apple juice, rapidly break down under the action of yeasts and form alcohol. Carbon dioxide gas is evolved here as in all *fermentation.*

Apparatus — Erlenmeyer flask, wide-mouth bottle, beaker.

Part I. Pour 25 cc. of thick molasses into 200 cc. of water and fill your Erlenmeyer flask with this solution nearly to the top. Break a small piece of yeast cake into a powder, drop the powder into the solution, stopper the flask, and connect it with the bottle as shown in the illustration. Fill the bottle half full of lime water. Allow the tube from the flask to dip below this lime water. Leave the apparatus in a warm place in the laboratory for at least twenty-four hours or until the next laboratory period. Perform Exercise 52 during the remainder of the hour.



20. THE FERMENTATION OF MOLASSES. — Note that the delivery tube does not dip beneath the stopper of the flask on the right which contains the fermented molasses, but that it does run nearly to the bottom of the bottle containing the lime water. The bottle containing the lime water should be closely covered with a cap of filter paper to partially prevent the carbon dioxide in the room from taking effect upon the lime water in the bottle.

Part II. At the beginning of the next laboratory period test the solution for carbon dioxide gas by noting the effect produced upon the lime water. The remainder of the solution is now, after fermentation, largely a mixture of impure alcohol, acetic acid, and probably some unchanged molasses. Our next step is to separate the alcohol as completely as possible from the mixture by a crude type of fractional distillation. We accomplish this by heating the flask containing the solutions and allowing the vapors formed to run across the delivery tube and down to a beaker placed below. Cool the tube by a wet cloth or sponge. Do not allow the mixture in the flask to boil over and run down across the tube. Avoid this by keeping the heat sufficiently gentle to allow only vapor to pass off. Alcohol boils below the temperature of water and, with care, a small amount of clear liquid can be obtained in the beaker. When 10–15 cc. have been secured, pour this clear liquid into a clean flask and repeat the process with this condensed liquid just obtained. After the second distillation see if you can obtain the odor of alcohol. Try to light the alcohol from the flame of your burner.

B

Describe the general action of fermentation as shown in this case.

Explain the principle of fractional distillation.

What are the common uses of alcohol?

Exercise 52. Formation of a Salt of Acetic Acid. Test for Acetic Acid. — Acetic acid, the acid of vinegar, forms several acetates or organic salts. One of the most typical of these compounds is known as ethyl acetate, a compound formed from acetic acid and alcohol. This compound is analogous to the essences of fruits and flowers. Here, as in other cases, the sense of smell proves of valuable assistance in recognizing the compound.

Apparatus — Test tubes.

Cautiously add a few drops of concentrated sulphuric acid to 4 cc. of acetic acid and alcohol mixed in equal parts in a test tube. Shake and warm gently. Smell. The odor is due to *ethyl acetate.* Smell of the ethyl acetate on the shelf and compare it with that which you have just formed. Test *vinegar* for the presence of acetic acid by repeating the last experiment, using vinegar in place of acetic acid. Note the smell.

Alcohol may be considered as an organic hydroxide. Acetic acid is a typical organic acid. Keeping this in mind and disregarding the action of the sulphuric acid, answer the following questions:

How does the action which produced the ethyl acetate here compare with the action by which sodium sulphate was produced from sodium hydroxide and sulphuric acid?

Give some of the common uses of acetic acid.

Exercise 53. Study of One of the Nutrients, the Carbohydrates: as shown in Sugar, Starch, and Cellulose. — In this experiment we are to consider, one great class of the nutrients, the *carbohydrates;* found in the sugar of sugar cane or of fruit; present in starch, such as the starch of potato; occurring in the cellulose which largely comprises the woody fiber of trees and vegetables. All the carbohydrates contain carbon combined with hydrogen and oxygen, present in the same proportions in which they exist in water. The formulæ for these bodies illustrates this fact. The formula for cane sugar is $C_{12}H_{22}O_{11}$; for glucose, $C_6H_{12}O_6$; for starch, $C_6H_{10}O_5$; for cellulose, $(C_6H_{10}O_5)n$. The " n " following the symbol for cellulose indicates that we are not quite sure how many times this group occurs in the molecule. In the following experiment we are to study the composition of three typical carbohydrates. We use *concentrated acid here. Take care.*

Apparatus — Test tubes.

Part I. Heat starch, wood, and sugar separately in test tubes. The wood may be a few broken unburned ends of matches. Note the sides of the tubes before any melting takes place. Then test for *solubility* by washing out the test tube.

Part II. To a small amount of fresh samples of wood and sugar in test tubes, add a few cc. of concentrated sulphuric acid. Note result and state what happens. Note the heat generated. Remember that sulphuric acid is hygroscopic. Look up hygroscopic if you have forgotten what it means.

B

What appeared on the sides of the tube after heating?

What appeared to be left after the sulphuric acid had ceased its hygroscopic action?

What three elements appear to be present in this class of nutrients?

Exercise 54. Soap-making. — The common household method of making soaps, used almost universally fifty years ago, was to boil soap fats, composed of the refuse fatty waste of the kitchen, with potash obtained from wood ashes. We may perform the experiment more easily in the laboratory by using castor oil in place of soap fats. Soap is largely a mixture of the salts of several organic acids (chiefly stearic and palmitic acids) existing in the ordinary soap fat of the kitchen. When soap fat is boiled with an alkali like sodium hydroxide the process is not unlike neutralization. The organic salts present break down in the process and produce glycerine (which is really an alcohol) and soap which, under these circumstances, is largely the sodium salt of the various acids. Soap, at first practically a liquid, is brought into a solid state by the addition of salt. This solidification takes place for the reason that organic compounds present in soap and soluble in water are largely insoluble in a salt solution. This insolubility of soap in salt water is shown every time that soap is used at the ocean. In sea water it will not go into solution or lather as in ordinary fresh water.

Apparatus — Evaporating dish. (*This should be of at least* 150 *cc. capacity, to allow for the quantities given here. If the dish has a less capacity than this divide your quantities in two.*) **Beaker.**

In an evaporating dish mix 25 cc. of castor oil and 25 cc. of sodium hydroxide solution (made up 3.5 gm. of solid to 25 cc. of water). Boil for 20 minutes, using a flame which *will not allow* the liquid to boil over. While boiling measure out 40 cc. of water in a beaker and make ready 20 gm. of salt on weighing paper. At the end of 20 minutes stop boiling, and, when the mixture is slightly cooled, add 40 cc. of water to the solution. Then bring it just to a boil again and add the 20 gm. of salt. Set the mass aside to cool until next day. Then skim off the *soap* and leave it to harden. *Note that the amounts given here will vary with different grades of castor oil.*

B

Discuss this formation of soap from the chemical standpoint, that is, state approximately what chemical changes occur.

NOTE. — Look forward into the next experiment and note, under the heading "Apparatus," that samples of wool and silk and cotton cloths or threads should be brought from home.

Exercise 55. Direct Dyeing. Dyeing by the Use of Picric Acid. — The dyeing of cloth is dependent upon the placing within the fibers of the textile of a colored substance which is insoluble in water. The fibers of cloth may be animal, like those of silk or wool, or vegetable like those of cotton and linen. Some animal fibers are filled tubes, the vegetable fibers are hollow tubes. This experiment is illustrative of direct dyeing. In it a substance contained in every fiber of the wool and silk used unites directly with the special dye (picric acid) used here. Cotton and linen do not naturally contain such substances in their fibers. As they hold nothing with which the dye can combine they cannot be dyed directly.

Apparatus — Small pieces of white flannel or light colored silk or white woolen yarn, small pieces of cotton cloth (*these should be brought from home*), **evaporating dish, flask, stirring rod.**

Part I. Dissolve 1 gm. of picric acid in 125 cc. of water in a flask. Pour one-half of the picric acid solution into your evaporating dish. Heat it for six or seven minutes, and immerse the wool or silk in it. Bring the liquid to a boil, then remove the material and rinse thoroughly.

Part II. Wash your evaporating dish thoroughly and repeat the experiment, using cotton instead of wool or silk. Note results.

B

What is the action of the picric acid?
Why did it dye the first material and not the second?
What is the structure of cloth?

Exercise 56. Indirect Dyeing. Dyeing by the Use of a Mordant Acting upon a Dye and Forming a Lake. — This experiment illustrates indirect dyeing, the common method. In this process a substance is precipitated within the fiber by the action of two external chemical agents. The dye added, which may be acid or basic, is acted upon by some mordant which precipitates a colored substance within the tube of the fiber. The colored precipitate so formed is called a lake. Lakes may be formed within the hollow tubes of vegetable fibers as well as within the filled tubes of animal fibers. This process, therefore, can be used with either animal or vegetable fibers.

Apparatus — Wool, silk, and cotton, as before, evaporating dish, flask, stirring rod, beaker.

Drop four or five drops of alizarin paste into 100 cc. of water in an evaporating dish and place on a ring stand to heat. Dissolve 2 gms. of iron acetate in a 100 cc. of water in your beaker. Pour 10 cc. of ammonia into 100 cc. of water in your flask. Immerse a piece of wool or silk and a piece of cotton, first in the iron acetate, then in the ammonia. Allow the pieces to soak a couple of minutes in both the ammonia and the iron acetate solution. Then place them in alizarin. Boil them in the alizarin solution. Finally rinse them thoroughly.

B

How does this action differ from the action in the last experiment?

Which of these processes seems to be the best all-round method?

Where do most of our dyes come from?

Explain briefly how dyeing is carried on.

Exercise 57. Study of Illuminating Gas. — The actual making of illuminating gas for commercial use requires a complicated apparatus. When coal is heated away from air, however, in such a fashion that the carbon it contains cannot oxidize to carbon dioxide and the hydrogen present oxidize to water, illuminating gas containing methane, carbon monoxide, hydrogen, etc., is formed. This experiment, beside showing the *production* of *illuminating gas*, illustrates the process of *destructive distillation.*

Apparatus — Ignition tube, large rubber tubing, pneumatic trough, burner.

Place in an ignition tube enough powdered bituminous coal to one-third fill the tube. Allow the powder to lie along the lower side of the tube with a free space above it. Attach wide rubber tubing to the glass tube, heat and collect the gas over the water. Test the collected gas, the bottle held mouth upward, with the flame. Collect a bottle of the ordinary gas from the gas tube on the desk by upward displacement, and compare this gas with the gas just made as regards the characteristic properties of each.

B

What are the properties of the illuminating gas made from coal?

In what other way can illuminating gas be manufactured?

Exercise 58. Preparation and use of Photographic Blue Paper. — For this experiment photographic printing frames and a few negatives, preferably of landscapes, are required. It is generally possible to borrow frames and negatives from pupils if the laboratory does not possess them. The blue paper made in this experiment breaks down under the influence of light in a fashion similar to the way in which our ordinary printing papers do. All photographic printing illustrates light reactions.

Apparatus — Test tubes, evaporating dish, printing frame, glass rod, cotton wool.

Part I. Mix 2 gms. of ammonium citrate of iron with 12 cc. of water in a test tube. Mix 2 gms. of potassium ferricyanide (red prussiate of potash) with 12 cc. of water. Do not handle either of the solids with your fingers. Use weighing paper altogether. Shake the test tubes until the substances have wholly dissolved. Then mix the two solutions together in a clean evaporating dish, stirring with a glass rod.

Take three sheets of common glazed block paper and give each a thorough coating on one side with the mixture, applying it by means of a bit of cotton wool. After coating the paper, handle it only by the sides and pin it up on the sides of a dark closet to dry until the next laboratory hour. Spend any time remaining of the period in thoroughly cleansing your glassware in preparation for the succeeding experiments.

Part II. Take a piece of the paper prepared last time, cut it to the size of the negative to be used and print a proof from one of the negatives in a printing frame. Place the frame in direct sunlight and leave it there until the picture shows clearly. While this is going on proceed with the next experiment. After printing remove the proof from the frame and wash in *hot* water.

B

What chemical agency have we used to bring about this change?

Look up the chemistry of photography and tell all you can about it in a short paragraph.

PRELIMINARY NOTE

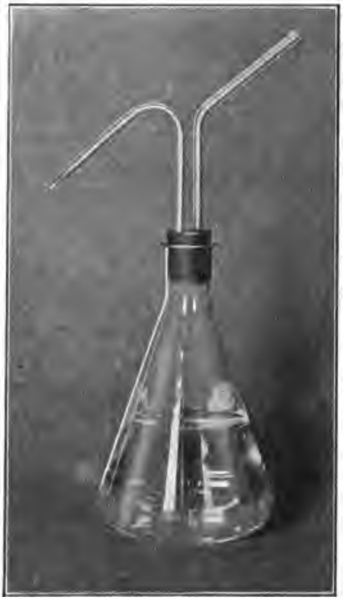
In the following experiments, 59, 60, 61, 62, 63, 64, we are to consider some of the properties of a few of the more important metals. In this consideration we shall try to distinguish certain properties which separate each metal from other metal. Every metallic element can be distinguished from every other metallic element by the use of tests similar to those given here.

In doing these experiments make no precipitations in graduates. Transfer all liquids to test tubes before testing them. In using the hydrogen sulphide generator in the course of the following experiments, be sure to turn off the gas immediately after using. Use a clean glass tube of your own each time you pass this gas into a solution.

You will find it to your advantage in these exercises to use a wash bottle like the one formed in the experiment on bending glass, Exercise 80, which follows. A wash bottle has two tubes: one, an acute bend, which extends to the bottom of the flask inside, has its outer end drawn into a tip; the other, an obtuse bend, reaches only to the bottom of the stopper. By blowing through the second bend we can force a stream of water from the first bend. This stream can be used to good advantage in transferring precipitates to filters.

Note that when a substance reacts with hydrogen sulphide gas, a sulphide of the substance is generally formed. For example, when copper nitrate reacts with hydrogen sulphide gas, copper sulphide is formed.

In recording these experiments enter your data in a systematic fashion. For example, in the following exercise on iron, and in the succeeding experiments, record each action with its results in as brief fashion as that which follows: "Passed hydrogen sulphide gas through the solution of iron sulphate: no precipitate fell."



21. Wash Bottle. — Note that the tube with the obtuse angle, on the right, extends only to the bottom of the stopper while the tube with the acute angle runs nearly to the bottom of the flask.

Be sure all windows of the hood are closed while the hydrogen sulphide generator is running.

Exercise 59. Study of Iron. Preparation of Ferrous and Ferric Salts. Oxidation and Reduction. — In this experiment we are to study ferrous and ferric iron, the two common forms of this metal. Several of the metals appear in two forms with different valences. Ferrous iron has a valence of two. Ferric iron has a valence of three. We are to make both a ferrous and a ferric salt here, the first by the direct action of an acid on the metal, the second by the oxidation of one form to the other.

There are a number of simple tests to be made in this experiment, so many that a class can scarcely finish in a laboratory hour unless all the reagents are at hand. Small amounts of each reagent used should be placed on every desk to do away with the time lost in obtaining them.

Enter your tests in the following fashion:

Ferrous chloride, Action with potassium ferrocyanide.
Action with potassium ferricyanide.

Apparatus — Test tubes.

PART I. TESTS TO DISTINGUISH FERROUS AND FERRIC SALTS. — Divide a small amount of a solution of ferric chloride between two test tubes and dilute with water until each test tube is one-third full. Divide a similar amount of a solution of ferrous chloride between two more test tubes and dilute as before. Either place the test tubes in such a position that they will not get confused, or mark them in some way. Into one of the tubes containing ferrous chloride, and into one of those containing ferric chloride, pour potassium ferricyanide. Compare the tests and record in a table as shown above. Into the other two tubes pour potassium ferrocyanide. Compare results and record. Note that one dark blue precipitate is formed by one reagent, and another similar precipitate by the other.

PART II. FORMATION OF FERROUS SULPHATE. — Place a small amount of soft iron wire or a small pinch of iron powder in a test tube and add enough dilute sulphuric acid to about one-fourth fill the tube. If necessary, heat to boiling to accelerate the action. Place the tube in the test-tube rack and allow the action to continue until the iron dissolves or action ceases. Ferrous sulphate is formed. Take a small portion of the solution, pour it into a watch glass and set it aside to crystallize. Take a second portion and test by adding a few drops of potassium ferrocyanide and potassium ferricyanide. Note any change and record. Go on with Parts III and IV while waiting for the iron to dissolve.

PART III. OXIDATION OF FERROUS CHLORIDE TO FERRIC CHLORIDE. — Take a small portion of ferrous chloride solution in a test tube, add a small amount of concentrated hydrochloric acid, then an equal amount of concentrated nitric acid. Heat just to boiling. Ferric chloride is formed by the oxidation of ferrous chlorides. Test with potassium ferricyanide and potassium ferrocyanide. Note any change and record. This is another example of oxidation produced by nitric acid.

PART IV. REDUCTION OF FERRIC CHLORIDE TO FERROUS CHLORIDE BY THE AID OF HYDROGEN. — To a small portion of fresh ferric chloride solution add a little dilute hydrochloric acid and either some soft iron wire or a small pinch of iron powder. Heat gently, if necessary, and note the gas formed. Hydrogen is a great reducing agent. Ferrous chloride is formed. When action ceases, test with potassium ferricyanide and potassium ferrocyanide. Ferric chloride should be reduced to ferrous chloride. Reduction takes place. If the test does not show ferrous iron at once, heat the mixture to boiling and test again.

B

How may we distinguish between ferrous and ferric salts?

Explain the oxidation of ferrous iron to ferric iron and the reduction of ferric iron to ferrous iron.

Give methods for making both ferrous and ferric salts.

What differences are shown by the endings, -ous and -ic?

Write the formulæ for ferrous and ferric chloride, for ferrous and ferric sulphate.

Explain oxidation and reduction in general.

Exercise 60. Study of Copper. — In the use of the blow-pipe remember that the essential thing is to keep up a steady blast of air. This may be done by filling the cheeks and breathing in through the nose, giving a steady draft. If the blow-pipes are common to the laboratory use your individual piece of rubber for a mouthpiece. Return the pipes when through using them.

Apparatus — Three test tubes, charcoal, blow-pipe.

To about 5 cc. of copper sulphate in a test tube add ammonia water until distinctly alkaline. Note the result. To 5 cc. more add hydrochloric acid. Note result. To 5 cc. more add hydrogen sulphide gas from the generator in the hood and note result. Filter into an evaporating dish and boil in nitric acid any precipitate formed, together with the part of the filter paper that holds it. *Do this last piece of work under the hood.* Mix equal weights of copper oxide and sodium carbonate on your piece of charcoal and reduce it with the blow-pipe, as shown in the illustration in Ex. 46.

B

How can you get a copper precipitate from a solution of a copper salt?

How can you put the precipitate back into solution again?

Write the reaction for the addition of hydrogen sulphide gas to copper sulphate.

Give a method for separating metallic copper from its compounds.

Exercise 61. Study of Silver.

Apparatus — Evaporating dish, test tubes, funnel, charcoal.

Fill your evaporating dish with water and put it on to boil. Take 5 cc. of silver nitrate solution and add ammonia. Take 5 cc. more and pass hydrogen sulphide gas through the solution. Take 10 cc. more and add hydrochloric acid. Note the character of the precipitate formed in each case. Filter each precipitate separately. Filter the precipitate formed by the hydrochloric acid in two parts. Pour ammonia through each filter paper and note carefully the action which takes place in the solution.

B

In this experiment what chemicals can you add to precipitate the silver from the solution of a silver salt?

Which chemical that we have used has precipitated nothing but silver?

How can you put the precipitate just mentioned back into solution again?

Write the reaction for each of these cases excepting the reactions with ammonia.

NOTE. — This experiment on the properties of lead is introduced before its normal place to allow time for the brief example of work in qualitative analysis introduced in the following experiments:

Exercise 62. Study of Lead.—Note: In this experiment, in addition to tests similar to those already used, we are to consider the solubility of a metal (lead) in the three acids chiefly used in the laboratory, nitric, sulphuric, and hydrochloric. Each acid used in this experiment is to be made up by adding 5 cc. of concentrated acid to 20 cc. of water.

Apparatus — Evaporating dish, test tubes, funnel.

Part I. Place a small piece of lead wire in each of three test tubes. Pour a few cc. of nitric acid in one. Pour a few cc. of hydrochloric acid in the second, and a few cc. of sulphuric acid in the third. Warm gently and observe the action.

Note carefully the appearance of the lead before and after it goes into the acids. Heat long enough to see whether or not solution is taking place, but do not try to get all of the lead into solution.

NOTE. — In this experiment, in order to give an opportunity for a consideration of problems of solubility we shall omit all tests with different reagents except the one with hydrochloric acid needed in the succeeding experiments.

Part II. Fill your evaporating dish with water and put it on to boil. Take 20 cc. of lead nitrate and add to it a few cc. of the hydrochloric acid on the desks. Take 5 cc. of lead nitrate and add a few cc. of the salt solution to it. Note precipitate formed. Filter the first precipitate. When completely filtered pour boiling hot water through the precipitate on the filter paper. Note results.

Part III. Place a little lead oxide on your charcoal and reduce it with your blow-pipe as in the preceding experiment.

B

Write the reactions for the salts of lead formed in Part I.

Was there any difference in the action of the three acids?

How can you get a lead precipitate from a solution of a lead salt?

How can you put it back into solution again?

Exercise 63. Qualitative Analysis of the Silver Group of Metals. — The analytical chemist depends, as has already been said, upon two types of analysis for his determinations: qualitative analysis, a procedure by which he can tell what elements are present in any compound or solution, and quantitative analysis, which tells him how much, by weight, of each element is present. The metals in qualitative analysis form naturally a series of groups whose boundaries are determined by the action of certain reagents. We have already seen under our study of chlorine and hydrochloric acid, of silver and of lead, that silver and lead compounds are precipitated by the action of a chloride. Three and only three of the metals, lead, silver, and mercurous mercury, are precipitated as insoluble chlorides in this way. These three are known as the silver group. By a series of precipitations we may separate each one of these metals individually from the others. The following experiments, preceded by an introductory exercise on mercurous mercury, show the qualitative analysis of the silver group.

Apparatus — Test tubes, funnels, filter paper, evaporating dish.

Part I. Take 10 cc. of a solution of mercurous nitrate and add hydrochloric acid. Note the result. Add ammonia. Note the change. We repeat this experiment in Part II in comparing the action of the reagents used on the different solutions.

Part II. Take 10 cc. respectively of solutions of silver nitrate, of lead nitrate, and mercurous nitrate ($AgNO_3$), $Pb(NO_3)_2$, $Hg(NO_3)$, and add hydrochloric acid to each. Note the results and record. Do not get the solutions confused with each other. Now treat the silver chloride and the mercurous chloride, formed in these test tubes, with ammonia. Follow this action by the addition of nitric acid to each test tube. Note results and record. Filter the precipitate of lead chloride and pour boiling hot water through the filter paper. Treat half of the solution of lead chloride thus formed with a few drops of concentrated sulphuric acid. Treat the other half with a little potassium chromate. Note the result and record. From these results answer the following questions:

B

What reagent precipitates all three of these chlorides?
How may we dissolve lead chloride?
How may we dissolve silver chloride?
What reagent causes a characteristic change in mercurous chloride?

Exercise 64. Continuation of the Analysis of the Silver Group. — Mix together 10 cc. of silver nitrate, of lead nitrate, and mercurous nitrate; add the reagent, just considered, which you found would precipitate all three of these metals and filter the precipitate obtained.

Add some substance which will dissolve out the lead chloride and leave silver chloride and mercurous chloride on the filter paper. Add to the filtrate chemicals which will test for a solution of lead chloride. Silver chloride and mercurous chloride are now upon the filter paper. Add a solution which will dissolve the silver and leave the mercury on the paper. Add a reagent which will reprecipitate the silver from the filtrate obtained.

B

Draw up a scheme for the systematic separation of these three metals of the silver group.

NOTE

A series of experiments, to be inserted at the most convenient point, follow. If time or apparatus is lacking for the performance of these exercises during the regular laboratory hour, they may be used as lecture-table experiments during recitation periods.

Exercise 65. Formation of Compounds (Oxides) of Iron and of Magnesium. — One of the most vexing problems which confronted the early discoverers in science concerned the weight of substances left after burning. They knew that when wood or coal burned the *visible* compounds (ashes) left behind were far lighter than the substances present in the beginning. It was a long time, however, before they realized that the *invisible* gases produced by burning, which passed off into the air, made the sum of the weights of the *visible* and *invisible* compounds, left after burning, greater than the weight of the fuel burned. One of the things that troubled them was the way in which metals acted when they were ignited. In this experiment we are to ignite two metals to form metallic oxides and, in so doing, consider the weights of these substances before and after burning. Note. — Use iron powder reduced by alcohol here.

Apparatus — Crucible, crucible triangle.

With some fine sand exactly counterpoise a crucible about one-third full of iron powder. Heat the iron for ten minutes, and, when cold, see how its new weight compares with the former one. Is it lighter or heavier? While the iron is cooling, light a strip of magnesium ribbon in the burner flame and examine the powder left after burning. Do not look directly at the flame of the burning metal, and hold the magnesium away from you over the sink. (Magnesium, by the way, is the flashlight powder of the photographer.)

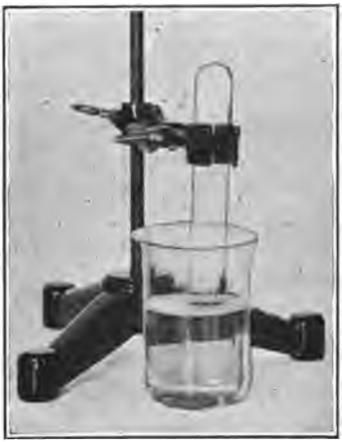
B

If there is an increase of weight in the first case, something has been added to the iron.

Where did it come from?

Evidently the magnesium has greatly changed its structure. How do you think this chemical change was brought about?

Exercise 66. Action of Sodium on Water. Sodium and Caustic Soda, Sodium Hydroxide. — In the following experiment we use sodium amalgam (which is sodium alloyed with mercury) to produce the important base, sodium hydroxide. Sodium is used in this form in order to make its action less energetic, and therefore safer, than it would otherwise be. The mercury used has no effect upon the result of the experiment. This reaction takes place between the sodium and the water.



22. ACTION OF SODIUM ON WATER. — The ignition tube is clamped upside down in a burette clamp. The beaker should contain enough water so that the tube can be readily inserted, but should not contain so much water that the gas evolved in the tube will cause the water in the beaker to overflow.

Be careful here, as in all experiments where mercury is used, to *keep the mercury away from gold rings.*

To get the tube full of gas during the laboratory period, it is necessary to get the apparatus working as soon as possible. If the amalgam given you does not seem to be producing gas enough, take more.

Apparatus — Evaporating dish, ignition tube, litmus and turmeric paper, watch glass.

Invert an ignition tube full of water in an evaporating dish two-thirds full of water, holding it there about one-half inch above the bottom of the dish by means of a burette clamp as shown in the illustration. Then take a lump of sodium amalgam and drop it below the tube, allowing the gas to rise into the tube. Test the water in the dish with red and blue litmus papers and with turmeric paper. *Make all possible tests while the gas is coming off.* Drop one drop of the liquid on a watch glass and let it dry. Take a platinum wire and dip it into the liquid. Hold the wire over the flame. Note the color of the flame. When the tube is full of gas remove the tube, keeping it mouth down, and apply a lighted match. Note: Filter off the water in the dish at the end and place the remaining mercury in the bottle labeled, "Mercury Residues." *Never* thrown any mercury away. *Never* let any get into the sink, as it ruins the pipes.

B

What gas is formed?

What is the solution left in the dish?

What is the flame test for sodium?

Is the density of metallic sodium (not sodium amalgam) greater or less than one?

Exercise 67. Distillation. — By distillation we may cleanse a liquid from solid impurities, or separate two liquids when one boils at a higher temperature than another. If we boil water contaminated by dirt (in boiling we are changing liquid water to gaseous water), pass the steam through a cooling tube and condense it in a beaker at the other end, as shown in the illustration, we shall obtain pure water; the dirt will be left behind. Such methods of distillation have proved of great value in manufacturing processes as well as in producing pure water from impure water. Distillation, as a mode of cleansing drinking water, is especially effective in times of emergency. Our ordinary water, however, needs some simpler and less expensive mode of purification. In consequence, filtering on a large scale is generally employed.

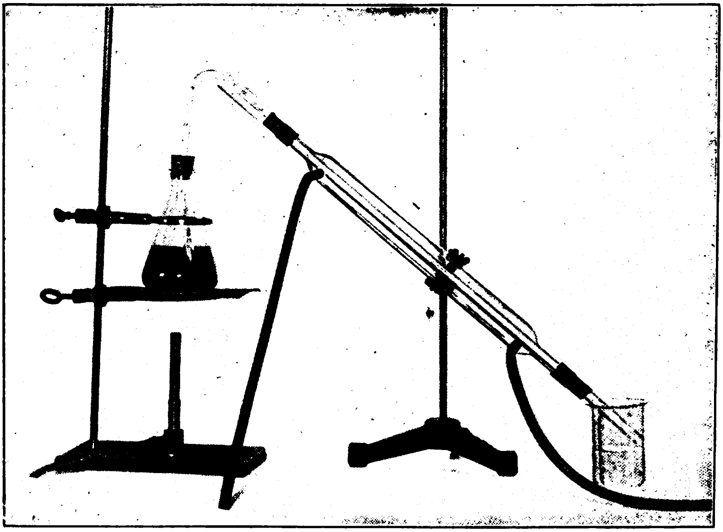
Apparatus — Condenser, flask, asbestos, beaker

Take a condenser like that shown in the illustration and place in it a solution colored with a few drops of copper sulphate solution (used to represent an impurity). Boil this, keeping the flask on asbestos. Place a beaker under the end of the condenser. Collect and examine the liquid which falls into the beaker.

B

What is the principle of distillation?

What is the value of distillation when the water supply is impure, or when ships are at sea?



23. Distillation. — The liquid in the flask, turning to gas, passes down the inner tube of the condenser (Liebig Condenser), is turned to liquid by cooling water, which flows through the outer tube or jacket, and runs as liquid into the beaker at the end. The cooling water enters by the lower rubber tube and passes out by the upper rubber tube. Any solid impurities present in the original liquid are left behind in the flask.

Exercise 68. Preparation of Hydrofluoric Acid. Etching on Glass. — *Hydrofluoric acid,* prepared in this experiment, requires the *greatest care* in handling. It must be kept *under the hood* during the whole period of the experiment. It *must not be inhaled.* It *must not be spilled* upon the hands. It is *poisonous and corrosive.* Some knowledge of etching and hydrofluoric acid is, however, too important to be wholly omitted. This experiment is a satisfactory lecture table experiment.

Apparatus — Small glass plate (about four inches square is a common size), lead dish, file.

Warm a glass plate over a radiator or by placing it on asbestos and heating the asbestos gently. Do not hold it over the direct flame where it is likely to crack. When warm, coat one side with a thin layer of paraffin. Take a file and carefully print the initials of your name on the coated side, making sure that the file cuts clearly through the wax to the glass below. Do not make the lines too narrow.

Take a crystal of fluorspar (calcium fluoride) and examine it carefully. Compare it with chlorides, bromides, and iodides that you have seen. Powder a little fluorspar in a mortar and weigh out from five to seven grams of the powder. Spread this quantity in a thin coating over the bottom of the lead dish.

THE REST OF THE WORK MUST BE DONE UNDER THE HOOD.

Place the lead dish in the hood at one side. Pour on to the powder enough concentrated sulphuric acid to make a paste about as thick as ordinary mucilage. *Carefully,* without disturbing the dish, place the glass plate over the mixture, wax side down. Leave the apparatus in the hood until the next day, then take the glass plate and remove the wax coating by drawing a heated wire across it. Do not try to melt the wax off with a flame. Take off the last portion of the paraffin by means of a filter paper wet with alcohol. The printed letters should be etched in the glass. Leave the lead dish with its contents in the hood to be removed in a separate waste jar.

B

What is the chemical action of the acid on the glass?

Where have you seen etched glass?

Can you suggest any way in which the etchings sold in the picture shops could be produced?

Exercise 69. Preparation of the so-called "Sodium Bicarbonate," Baking Soda, Sodium Hydrogen Carbonate. An Illustration of the Solvay Process. — In this experiment we are to illustrate the manufacture of the widely used salt, common baking soda, sodium hydrogen carbonate. We shall not carry the process to completion and prepare normal sodium carbonate. Read the description of both the Leblanc and the Solvay processes in the text-book.

Apparatus — Flask, generator, beaker.

Dissolve ten grams of ammonium carbonate in 100 cc. of strong ammonium hydroxide (common ammonia) in a flask. The solid must be ground to a powder in a mortar, and the solution heated until all the solid disappears. Allow to cool. While the cooling proceeds, weigh out twenty grams of salt (sodium chloride). Add this salt in four portions, shaking constantly. This amount should saturate the solution. Remember that in a saturated solution the dissolving liquid has taken up all the solid that will dissolve. In this case quite a little solid residue should be left after saturation has been obtained. The liquid left at the end of this action represents the first stage of the manufacturing process in which a salt solution is saturated with ammonia. When the solution appears to be saturated, allow the solids to settle and filter. A clear liquid should pass through as a filtrate. Prepare a carbon dioxide generator and make carbon dioxide as in Exercise 21, using marble chips and hydrochloric acid. Pass the gas into the solution until the end of the hour or until a layer of solid has covered the bottom of the flask. Place away until the next laboratory hour. The action just performed represents the second stage in the manufacturing process. Take the layer which has deposited, filter it off, allow it to dry and examine it. This process may be carried on during the next two laboratory periods. Compare the solid with the sodium hydrogen carbonate on the desk as regards its solubility, appearance, and action with hydrochloric acid. Note that the general appearance of the solid is quite different from that of either the salt or ammonium carbonate with which you began.

B

Describe the Solvay process briefly.

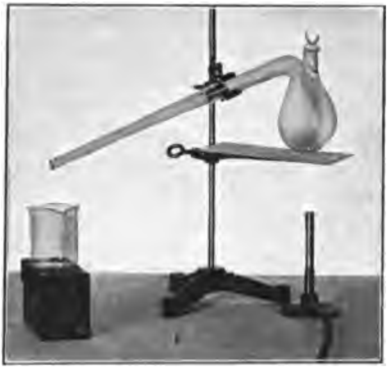
Distinguish between the two carbonates of sodium and state some of the purposes for which each is used.

Exercise 70. Preparation of Nitric Acid. — Nitric acid is one of the widely used commercial acids. It acts as a solvent of the metals. Nitrates, the salts of this acid, may be formed by direct solution in hot or cold nitric acid. The usual method of preparing this acid is to act upon either sodium or potassium nitrate with concentrated sulphuric acid. To effect the reaction the mixture is heated in a retort, whereupon the nitric acid distills over. Nitric acid itself is colorless and the red-brown fumes which appear in its preparation are due to the formation of nitrogen peroxide gas. *Great care should be taken in handling this concentrated acid. It is very corrosive.*

Do this under a hood.

Apparatus — Retort, asbestos, beaker.

Place twenty grams of sodium nitrate, which has been *gently* ground to a fine powder in a mortar, in a glass retort. Pour it in by means of weighing paper. Add, through the tubulure, 10 cc. of concentrated sulphuric acid. Be careful not to spill a drop. Place the retort on asbestos as shown in the illustration and heat gently, placing the beaker at the end of the retort tube or placing a wide-mouth bottle over the neck of the retort. Cool the neck of the retort from time to time with a wet sponge, allowing the water used to flow into a pneumatic trough. When a small amount of the gas has condensed to liquid acid, dilute with water by pouring the acid obtained into a beaker nearly full of water. Use the test for nitrates, already performed, with a part of the solution just obtained. If time permits form sodium nitrate, ammonium nitrate and potassium nitrate by neutralizing a few cubic centimeters of this solution with dilute solutions of sodium hydroxide, ammonium hydroxide and potassium hydroxide. Follow the methods given in Exercise 10.



24. MAKING NITRIC ACID. — The *retort* shown in the illustration is attached to the ring stand by a clamp so that its *bulb* rests on an asbestos resting on a ring. The narrow *neck* of the retort is at the left of the picture. If this experiment is done in a good hood the neck may be left open as here, but it is probably better in any case to let a wide-mouth bottle rest directly against the neck of the retort instead of using a beaker. The glass-stoppered *tubulure* of the retort appears just above its bulb The neck of the retort may be cooled with a band of wet filter paper, instead of with a sponge.

B

What are the properties of nitric acid?

Write the equation for this action.

Exercise 71. Preparation of the Pigment Chrome Yellow. Reduction of Potassium Bichromate to Potassium Chromate. Oxidation of Potassium Chromate to Potassium Bichromate. — The pigment chrome yellow, the basis of a large amount of yellow paint, is a chromate salt of lead. Its symbol is $PbCrO_4$, and its molecule contains, beside lead and oxygen, an atom of the element chromium whose symbol is Cr. Chromium is a gray metal, never found free and occurring chiefly in chrome iron ore ($FeCr_2O_4$). This metal is used to harden steel for safes and armor plate, projectiles, and crushing machinery. The steel made from it is called chrome steel. The three chief compounds of chromium beside lead chromate are chrome alum ($K_2Cr_2(SO_4)_4 \cdot 24H_2O$), potassium bichromate (dichromate) ($K_2Cr_2O_7$), and potassium chromate (K_2CrO_4). Both chrome alum and potassium bichromate are used largely in tanning and dyeing. The chromate and bichromate of potassium exhibit one interesting peculiarity. The chromate is oxidized to the bichromate by the use of acids. The bichromate is reduced to the chromate by the use of alkalies.

Apparatus — Test tubes.

Prepare dilute solutions of potassium chromate, potassium bichromate, and lead nitrate. Mark your test tubes in such a way as to know just what you have in each.

Part I. Pour one-half the solution of potassium chromate into the lead nitrate solution. Filter and examine the precipitate of *chrome yellow*, lead chromate, formed. While filtering proceed with Part II of this experiment.

Part II. Pour one half of your solution of potassium bichromate into a separate test tube. This leaves one half of each of the two original solutions to serve for comparison. To the portion of the solution just poured out add a small amount of concentrated hydrochloric acid. Note changes in color and compare with original solutions. Now to this same solution add potassium hydroxide until the solution becomes alkaline as shown by litmus paper or until the color changes. Note changes. Compare with the original solutions and record results.

B

Discuss the production of chrome yellow and explain the difference between a pigment and a dye.

Define oxidation and reduction.

Explain the method of producing oxidation and reduction in this especial case.

Exercise 72. Electrolysis of Water. — In setting up individual apparatus for this experiment, clamp two inverted ignition tubes or small test tubes side by side with burette clamps, in a similar fashion to the adjustment of the tube in the illustration shown in Exercise 66. The current should be introduced through platinum electrodes, which are run up inside the tubes. In general the student's apparatus should resemble the simple lecture table apparatus for electrolysis mentioned in the appendix.

Apparatus — Two test tubes, one large beaker, platinum electrodes.

Fill the beaker nearly to the brim with water, and invert two test tubes full of water over it in such a fashion that the mouths of the tubes will be a couple of inches below the surface. Clamp the burettes in place. Carefully push the platinum electrodes under the tubes, one to each. Note carefully which is the anode and which is the cathode. Then lower the water in the beaker to a point just above the mouth of the tubes. This excess water may be removed by a medicine dropper. It is taken out to prevent an overflow. Add a few drops of sulphuric acid to the water in the beaker. This acid need not be considered in observing the chemical changes which occur. Consider it simply as assisting the action. When all these preliminaries have been completed, connect the current. Observe the gases rising within the tubes. When either of the tubes is full of gas, cut off the current, take a centimeter rule and measure the length of the column in each tube. Notice whether the greater volume of gas is liberated at the anode or at the cathode. After measuring, remove the tubes cautiously, noting carefully which is which, hold them mouth down and test each one, first with a glowing splinter, then with a lighted match. It may be that only one gas will test satisfactorily when this is done. If both do not give positive tests refill the tubes again and test again. This time use the lighted match first and the glowing splinter second. Record all your results.

B

What gas is liberated at the anode?

What gas is liberated at the cathode?

Which is the anion and which the cation in this case?

What were the relative amounts of gas in the tubes? Which has the greater volume? Which the smaller?

What does this result indicate concerning the relative volumes of the gases which compose water?

Exercise 73. Borax Bead Tests.

Enamels and glazes, like the artificial gems which imitate the ruby, the sapphire, or the turquoise, commonly contain borax. The different colors of "paste jewels," like the various tints of enamels, come largely from the property which borax possesses of receiving a color from various metallic compounds. This property furnishes a valuable test for the metals in qualitative analysis. This test is illustrated in the following experiment. General solutions of each of the compounds used in this experiment should be provided for the use of the class.

Apparatus — Platinum wire sealed in a glass rod, test tubes.

Loop one end of a platinum wire into a small oval, heat it and dip it into powdered borax. Melt the powder in the flame, adding borax as necessary, until you have a small transparent bead filling the ring. Prepare in test tubes solutions of cobalt nitrate, chrome alum, nickel nitrate, copper sulphate, and iron chloride. Make these solutions by pouring a small quantity of the solution provided into the bottom of your test tube and filling the tube half full of water. Take a small portion of powdered manganese dioxide and place it beside your solutions on a weighing paper. Prepare a diagram like the following:

	Oxidizing Flame		Reducing Flame	
Metal.	*Color hot*	*Color cold.*	*Color hot.*	*Color cold.*
Cobalt				
Chromium				
Nickel				
Copper				
Iron				
Manganese				

Dip the borax bead in the cobalt nitrate and hold it first in the outer (or oxidizing) flame. Examine it when both hot and cold. It may be necessary to dip it into the solution several times to get the color. Repeat, holding the bead in the inner (or reducing) flame. Examine when hot and cold as before. Enter your results.

Remove the colored bead by heating it red hot and then plunging it instantly into cold water. This will loosen it so that it can be readily removed. Form a new bead and then proceed with the rest of the experiment.

Repeat the test performed with cobalt nitrate with each of the other solutions and with manganese dioxide, using a very small amount of the solid. If the bead turns black, add a little more borax. Enter your results as you obtain them. The results shown in the diagram will take the place of the usual questions under B.

Exercise 74. Analysis of Hydrochloric Acid Gas. — In this experiment we are to decompose hydrochloric acid gas by the action of sodium (used here in the form of sodium amalgam). From the volumes of gas before and after the reaction we may determine the relative volumes of the hydrogen and chlorine in the compound gas. The sodium unites with one part of the hydrochloric acid, forming a solid which appears on the side of the tube, while the other part of the gaseous acid is left free.

Apparatus — Ignition tube, generator for producing hydrochloric acid gas, pneumatic trough.

Generate the Gas Under the Hood.

Take a small piece of sodium amalgam about the size of a small pea on weighing paper. Take a clean, dry, ignition tube and get a cork stopper which fits it well. Generate hydrochloric acid gas in the same fashion as in Experiment 17, and dry the gas by means of a calcium chloride tube. Fill your ignition tube full of the gas by upward displacement. You can tell when it is full by the change of color of blue litmus paper held near the mouth of the tube. After the tube has been filled with the gas, reverse it quickly and drop in the piece of sodium amalgam. *Immediately* stopper the tube tightly with the cork stopper. Shake the tube (without allowing any of the gas to escape) for a couple of minutes, and then uncork it mouth downward under water. After water ceases to rise, raise the mouth of the tube till it is just below the surface of the water, and measure on the outside the amount of water which has risen into the tube. After noting the comparative amounts of water and gas, let the water escape and test the gas remaining in the tube with a match, holding the mouth of the tube down. Be careful not to allow any extra gas generated from the combination of sodium with the water in the trough to enter the receiving tube. To avoid this you should move the tube away from the sodium as soon as you unstopper it, keeping the mouth of the tube under water constantly. Examine the sides of the tube.

B

Hydrochloric acid gas is composed of two elements, hydrogen and chlorine. With which of these elements did the sodium unite?

What were the proportions of gas and water in the tube?

What became of the chlorine?

What did the flame test show with regard to the gas at the top of the tube?

Note.—The following experiments are quantitative and involve the use of the balance. In each of them we are to determine quantities of matter present. Exact weighing and careful manipulation are essentials for success. Read carefully the directions at the heading of the following experiment and apply the directions given there. These experiments will probably take several laboratory hours for their proper performance. At the end of each period of work put your materials and apparatus carefully away. Mark them, "Reserve."

Exercise 75. Determination of Water of Crystallization in Barium Chloride. — Note: In all experiments where weights are to be taken, never throw away the substance you are weighing until after your notes have been corrected. In all such experiments use the same balance for all your weighing, and enter your weights in your notebook as soon as they are determined.

This experiment, like the other quantitative exercises that follow, determines *what amounts* of certain elements are present, rather than *what elements* are present. The greatest care must be taken in all quantitative work to use the utmost cleanliness and care and never to lose a single particle of the material. Slight errors in manipulation make large ones in result.

Many substances used commercially have what is called a wet weight and a dry weight. Here the wet weight of the barium chloride used is the weight of the salt plus the weight of the water of crystallization which has joined with it.' The dry weight of the barium chloride is the weight of the salt after all the water has been driven out of it. It is evidently an important matter for those who buy substances such as this to know how much of the dry chemical they are buying, and how much water may be present in any given amount which they may purchase. The following experiment shows how this may be determined.

The water of crystallization in barium chloride is driven off by heat, leaving the dry or anhydrous salt. In consequence, to determine the amount of water present, we heat to constant weight a weighed quantity of the crystallized salt in an uncovered crucible. The loss will be the water of crystallization. The expression "*constant weight*" used here means the fixed weight obtained when all the changes which can take place through the action of ordinary heat have occurred.

Apparatus — Crucible.

Part I. Enter in your note-book the following data. Record each weight in its proper place as soon as you take it. Make all calculations in your note-book, and use no loose paper.

1. Weight of crucible alone.
2. Weight of crucible and crystallized barium chloride.
3. Weight of crystallized barium chloride alone. (This is to be found by subtracting No. 1 from No. 2.)

Weigh accurately a clean, dry crucible. Weigh in the crucible from 1.5 to 3 gms. (not over 3 gms.) of barium chloride, noting the exact amount taken. Record the results in the fashion shown in the data. Express weights in grams and fractions of grams, as 2.55 gms., not 2 gm. 5 dgm. 5 cgm. Cover your crucible and put it on one side until everything else is prepared. Then proceed with Part II.

Part II. Stand the uncovered crucible on a pipe clay triangle, and heat with a low flame for about two minutes, gradually increasing the heat until the tip of the blue cone is just below the bottom of the crucible. While heating take a fresh sheet of note-paper and do the problems in Part III. Continue the intense heat for fifteen minutes, then gradually lower the flame until it just touches the bottom of the crucible. Move the flame to and fro under the crucible until all danger of cracking disappears. Stand the crucible on asbestos. Weigh when cool enough to take in the hand. Record the results, thus:

4. Weight of crucible and crystallized barium chloride (No. 2 repeated).
5. Weight of crucible and dry barium chloride.
6. Weight of water of crystallization. (This is found by subtracting No. 5 from No. 4.)
7. Percentage of water of crystallization present in this sample of barium chloride.

To be sure that all water has been driven off, heat strongly for five minutes more, then cool and weigh as before. If there is any appreciable difference between these two weights repeat until the last two weights are constant.

B

If there is no change upon the second heating, calculate the per cent of water lost and the total number of molecules of water of crystallization per molecule of barium chloride, from the data given above, Do this by the method shown in the model problems below.

Part III. — Note: In solving these problems consult the first model problem given below.

1. Calculate the per cent of water of crystallization in a sample of barium chloride, 1.5 gms. of which lost 0.22 gm. on heating.
2. A student found that 2 gms. of barium chloride lost, on heating, 0.295 gm. What per cent of the sample was water of crystallization?
3. A sample of copper sulphate weighing 3.6 gms. lost .4 gm. on heating 10 minutes, and .7 gm. more on heating 20 minutes. What per cent was lost in each case, and what was the total per cent of crystallization (water of crystallization) in the salt?

First model problem on percentage of water of crystallization:

17 gms. of a given salt lost 3.5 gms. on heating. What was the per cent of water of crystallization lost?

$$\frac{3.5\text{ gms.} = \text{amount lost}}{17\text{ gms.} = \text{whole amount}} \times 100 = \%\text{ lost} \qquad \begin{array}{r} 17\,)\,3.5\ \ (\,.205 \\ \underline{3\,4} \\ 100 \\ \underline{85} \\ 15 \end{array}$$

The water was .205 of the whole amount. The percentage of water lost is found by multiplying by 100, which gives $.205 \times 100 = 20.5\%$ water of crystallization lost.

Second model problem to determine the total number of molecules of water of crystallization per molecule of a given salt:

5 gms. of crystalline sodium carbonate (Na_2CO_3) after heating leaves 1.85 gms. anhydrous Na_2CO_3. Find the number of molecules of water per molecule of sodium carbonate.

In the text, in the chapter on chemical arithmetic, you will find this proportion:

Molecular weight of whole: Molecular weight of part required :: Whole weight (in grams or pounds): Weight of part required (in grams or pounds). We shall use this proportion here.

The molecular weight of Na_2CO_3 anhydrous = 106.

$$Na_2 = 46, C = 12, O_3 = 48$$

Inserting the known values in the proportion given above,

$x : 106 :: 5$ gms.: 1.85 gms.

$1.85\,x = 530$ (106×5).

$x = 286$, whole molecular weight.

$286 - 106 = 180$, total molecular weight of water present.

Molecular weight $H_2O = 18$.

$H_2 = 2, O = 16$.

$180 \div 18 = 10$ molecules H_2O.

Na_2CO_3, 10 H_2O = symbol.

1. After heating 8.2 gms. of crystallized magnesium sulphate thoroughly, 4. gms. of the dry salt was left behind. Calculate the number of molecules of water of crystallization per molecule of magnesium sulphate.

$$Mg. = 24, S = 32, O = 16, H = 1.$$

Exercise 76. Determination of the Weight of Hydrogen Replaced by Zinc. — This experiment illustrates a method by which we can determine the relative weights with which zinc and hydrogen enter into combination with the sulphate radical. The action proceeds in accordance with the equation.

$$Zn + H_2SO_4 = ZnSO_4 + H_2.$$

The weights of the sulphuric acid and of the zinc sulphate are not necessary factors in the solution of the problem. The weights of zinc and hydrogen are necessary. The weight of the first is taken by the use of the balance. The weight of the second is obtained by finding the volume of hydrogen evolved in the reaction, and then calculating the weight of this volume from our knowledge that one liter of hydrogen weighs .09 gm. The method of computing the weight of hydrogen is shown in the model problem below. Knowing the weights of zinc and hydrogen we may find their comparative weights by dividing the weight of zinc by the weight of hydrogen, thus: $\frac{\text{weight of zinc}}{\text{weight of hydrogen}}$. There are two atoms of hydrogen, however, involved in this reaction. Zinc is bivalent and displaces two atoms of hydrogen. Therefore, to find how much greater the combining weight of the zinc atom may be than the corresponding weight of the hydrogen atom, we must divide the answer just obtained by two. This gives zinc's combining weight with regard to one atom of hydrogen, the standard.

This exercise illustrates a method similar to those used for determining the atomic weight of an element. It is *not* a determination of the atomic weight of zinc. Determinations of atomic weights are only made by years of labor on the part of thorough investigators.

NOTE:— The zinc used in this experiment is best obtained in the form of chemically pure zinc wire. To secure a weight of about one decigram it is well for the instructor to cut off several pieces first, weigh them and measure their length. A standard length of wire can then be obtained which can be given to students and which will weigh very close to a decigram. This will prevent the great loss of time resulting from experimenting with samples which are too heavy.

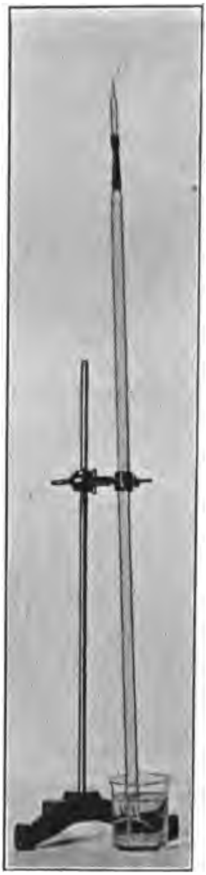
Enter the following data in your note-book before beginning the experiment.

1. Weight of zinc =
2. Number of cubic centimeters of hydrogen =
3. Barometric pressure =
4. Temperature of room =
5. Number of cubic centimeters of hydrogen at N.T.P. =
6. Weight of hydrogen =
7. Ratio of weight of zinc to weight of hydrogen =

Apparatus — The apparatus to be used is shown in the illustration.

It consists of an inverted burette secured by a clamp to a ring stand and placed over the top of a thistle tube whose tube has been cut short. The mouth of the burette and the thistle funnel are placed under water in a beaker. The piece of zinc wire, when placed inside the thistle tube and treated with acid, acts on the acid, evolving hydrogen. This hydrogen rises into the burette, and is collected there. The purpose of the thistle tube is to direct the bubbles of gas into the burette. Without it some bubbles are likely to escape. The fact that the burette is inverted makes it necessary to determine the unmarked volume at the bottom of the burette. This may be done by filling the burette (placed in its normal position) with water to the lowest mark, and allowing the liquid to run into a graduate.

Take a small amount of pure zinc wire (about 1 decigram) and place it under the funnel in the beaker. Be sure it is wholly under the funnel before proceeding. Arrange the apparatus like the model shown in the illustration. Place the thistle funnel in the beaker. The beaker is filled with water to a point just above the neck of the thistle funnel. Fill the graduated tube with water, close it with your thumb, invert it over the neck of the funnel and fasten it firmly with a clamp. Then, by means of a medicine dropper, take out the water until the liquid has nearly sunk to the level of the opening of the burette. This is done because the liquid displaced by the hydrogen must find room to escape into the beaker without overflowing onto the desk. By means of the medicine dropper introduce *concentrated* sulphuric acid into the water in the beaker until brisk chemical action sets in. Allow the action to continue until the zinc completely disappears. While the action is going on take a fresh sheet of paper and do the problems under Part B. After the apparatus has cooled to the temperature of the room, read the volume of the gas and calculate its volume at N.T.P. After calculating this, find the relative weights of zinc



25. Apparatus used in Exercise 76.

and hydrogen as shown below. Make all calculations in your notebook. Consult the text-book, pages 378–382.

PART B

In doing these problems refer to the model problem given below.

1. We have 280 cc. of nitrogen collected at 80 cm. of mercury pressure, and 40° C. Reduce to N.T.P.

2. We have 300 cc. of oxygen collected at 90 cm. of mercury pressure, and 50° C. Reduce to 80 cm. and 20° C.

3. What is Absolute Zero?

4. How much does any true gas expand or contract for a rise or fall of one degree Centigrade?

5. Model problem on Boyle's and Charles' Law:

We have 100 cc. of oxygen at 780 mm. and 27° C. How much would there be at N.T.P.?

$T : T' :: V : V'.$

$273 : \{(273 + 27) = 300\} :: x \text{ cc.} : 100 \text{ cc.}$

$300x = 27300$

$x = 91$ cc. at 0° C. and 780 mm. pressure.

Substitute the value of x (91 cc.) in the following proportion:

$P : P' :: V' : V.$

$760 : 780 :: 91 \text{ cc.} : x \text{ cc.}$

$760x = 70980.$

$x = 93.7$ cc. at N.T.P. or at 760 mm. and 0° C.

PART C

There are four steps in the calculation of the problem involved in the experiment just performed, all of which proceed from the two factors of the volume of hydrogen and the weight of zinc.

1. Reduce the volume of hydrogen (obtained by you experimentally) to N.T.P. as shown above. This must be done because the known weight of a liter of hydrogen .09 gm. is measured at N.T.P.

2. Find the weight of hydrogen by means of the following proportion. Let the volume of hydrogen (as corrected to N.T.P.) be represented by the letter "*n*."

Weight of a liter of hydrogen (.09 gm.) : Unknown weight of volume of hydrogen (x) :: Volume of 1 liter (1000 cc.) : Corrected volume of hydrogen (n).

$.09 : x :: 1000 : n.$

$1000x = .09 \times n.$ $x =$ the weight of hydrogen.

3. Divide the weight of zinc by the weight of hydrogen to find the combining weights as used here with two atoms of hydrogen.

4. Divide this answer by 2 to determine the relative combining weights of one atom of zinc and one atom of hydrogen.

Exercise 77. Determination of the Weight of a Liter of Oxygen.— In this experiment, as in the last, we are to find the weight of a measured volume of gas. Similar methods are used in both cases. In our work with hydrogen we generated the gas by the action of zinc on sulphuric acid. The weight of the hydrogen obtained was compared with the weight of the zinc which it replaced. In this experiment we are to generate oxygen from a known weight of potassium chlorate and manganese dioxide, and determine the weight of the volume obtained. The volume of oxygen, taken at room temperature and pressure, is reduced to normal temperature and pressure as the volume of hydrogen was reduced in the last experiment. The action proceeds as in the following equation:

$$KClO_3 = KCl + 3O$$

We can check our results experimentally in this exercise, as the loss in weight of the mixture heated gives the weight of the volume produced. From this knowledge of the weight of a part of a liter we can find the weight of a whole liter of oxygen. We shall afterwards compare it with the established weight of this quantity of the gas. One liter of oxygen at 0° and 760 mm. pressure weighs 1.429.

Apparatus — Long ignition tube closed at one end, glass wool, rubber stopper, half-bend glass tube, delivery tube, large pneumatic trough, 500 cc. graduate, glass plate.

Weigh out not less than 1.30 gms., and not more than 1.50 gms. of *dry* powdered potassium chlorate. Weigh out not less than 1 gm. and not more than 1.1 gms. of *dry* manganese dioxide. Mix the *powdered material*, thoroughly on weighing paper instead of in a mortar. (If they are ground together in the ordinary mortar too much of the solid is likely to stick to the sides of the mortar.) Place the mixture in the ignition tube, taking care that it does not stick on the upper sides of the tube, and insert a loose plug of glass wool above the powder. A rubber stopper with an oblique bend of delivery tubing should be prepared. With this close the tube tightly. Weigh the tube with its stopper, plug and half bend as accurately as possible. Using the apparatus just weighed, set up an apparatus for generating oxygen similar to that used in Exercise 5, excepting that the delivery tube is made of two parts, the oblique bend already mentioned, and a second tube connected tightly to the oblique bend by rubber connectors. The small pneumatic trough is replaced by a large trough, and the small bottle is replaced by a graduate of 500 cc. capacity which has been filled with water and inverted as the bottle was. This graduate may be secured by a burette clamp. Shake the ignition tube carefully in such a way as to spread the mixture over the bottom of the tube in a thin layer which may extend as far as, but no farther than, the middle of the tube. When these preparations are completed, heat and generate oxygen until it seems certain that protracted heating will give no more of the gas. Heat every part of the mixture thoroughly. After the action is completed allow the apparatus to cool to the temperature of the room, and read the volume of oxygen obtained. Reduce this volume to N.T.P. by the method shown in the last experiment. While the gas is cooling weigh the tube with every part adjusted in precisely the same fashion as before. Do this weighing with the utmost care. The weight thus obtained should be the weight of the volume of oxygen just calculated. From these factors figure the weight of a liter of oxygen according to the results of this experiment and compare the answer with the actual weight of a liter of the gas.

Enter the results obtained in the following data:

Weight of tube and contents before heating =

Weight of tube and contents after heating =

Weight of oxygen driven off =

Volume of oxygen obtained at room temperature and pressure =

Volume of oxygen at 0° and 760 mm. pressure =

Weight of 1000 cc. of oxygen according to the results of this experiment =

B

The calculation of results proceeds by four steps.

1. To obtain the weight of the oxygen used. This is found by subtracting the weight of the tube and its contents after heating from the weight of the tube and its contents before heating.

2. To obtain the volume of oxygen at N.T.P. This is calculated from the volume of oxygen experimentally obtained, by the methods illustrated in the preceding experiment.

3. To obtain the weight of oxygen in a liter according to the results of this experiment. This may be done by the following proportion.

Weight of oxygen used : Number of cc. of oxygen obtained :: Weight of one liter of oxygen : 1000 cc.

4. Comparison of the weight obtained experimentally with the weight of a liter of oxygen as given in the introduction to this experiment.

Exercise 78. Determination of the Combining Weights of Copper and Oxygen. — In this experiment we are to determine the combining weights of copper and oxygen by weighing a given quantity of copper oxide, weighing the tube which contains it with its contents, passing hydrogen over the copper oxide until all its oxygen, having combined with hydrogen to form water, has been driven off, and then weighing the containing tube again. The loss gives the weight of oxygen which was driven off. Subtracting this weight of the oxygen from the weight of the compound gives the weight of copper which remains.

Copper oxide has the formula CuO. Both copper and oxygen are bivalent. One atom of oxygen unites with one atom of copper. In consequence the weights obtained from the experiment should give the ratio of combining weights directly. In this, as in the preceding exercise, we are merely illustrating a method by which combining weights may be obtained. We are *not* conducting a research.

Apparatus — Hard glass tube 10 cm. long, open at both ends and fitted with stoppers which have glass tubes extending from each end. Generator like that shown in the illustration for Experiment 8.

Some points should be noted here. The hydrogen evolved from the generator on the left should be dried by passing through a calcium chloride tube similar to that shown in the illustration for Experiment 8. Two twists of copper wire should be placed around the piece of tubing which holds the copper oxide near the ends. These tend to prevent the burning of the stoppers. The tube containing the copper oxide is slightly inclined downward at the end away from the generator.

In this experiment you are to generate hydrogen and heat a tube through which the gas passes. Make sure by tests that the hydrogen is free from any mixture of oxygen before heating the tube. Have the generator some distance from the tube containing the copper oxide. Do not bend over the apparatus.

Weigh 10 gms. of copper oxide accurately, and put it in the combustion tube, then weigh carefully the combustion tube containing the copper oxide with its stoppers and tubing.

Heat the copper oxide. Generate hydrogen as in Experiment 8, and dry it by a drying tube. Pass the dry hydrogen in a slow stream over the hot copper oxide until the black color entirely disappears. Let the tube cool. Weigh the tube as at the beginning of the experiment. Repeat, passing more gas over the copper oxide until a constant weight is obtained. The loss of weight is due to the removal of oxygen.

Fill out the following data:

1. Weight of copper oxide =
2. Weight of combustion tube and copper oxide before heating =
3. Weight of combustion tube and copper oxide after heating =
4. Weight of oxygen which was driven off. (No. 2 – No. 3) =.
5. Weight of copper left behind. (No. 1 – No. 4) =.
6. Combining weights of copper and oxygen =

Exercise 79. A Chemical Equivalent Obtained by Displacement. — A remarkable phenomenon often takes place when a metal in its elementary form is placed in a solution of a compound of another metal. In such cases the compound in solution releases its metallic part, which comes out of solution, while the other metal, which entered as an element, goes into combination with the non-metallic part of the compound, producing a second and different combination. There are various laws which govern these interchanges, but we have scarcely space to consider them in detail here. It is sufficient to say that when we compare the weights of the metals which so strangely appear to exchange places, they seem closely related to those chemical equivalents which are discussed on pages 246 and 247 of the text-book. In the case before us we are to compare zinc and copper by placing a rod of pure elementary zinc in a solution of the compound, copper sulphate. We shall find that the zinc disappears as an element and combines with the sulphate radical to form zinc sulphate, while enough copper comes out of combination to equalize the amount of zinc which enters into combination. In other words, an equivalent amount of elementary copper precipitates to correspond to the amount of zinc which enters into the solution (in compound form). We weigh the elementary zinc before the action takes place. We dry and weigh the elementary copper which is left at the end of the experiment.

From these two figures we may find how much copper by weight is replaced in copper sulphate by a given amount of zinc.

Apparatus — Beaker.

Copy the following data into your note book and enter each weight when taken.

Data

Weight of zinc =
Weight of copper and filter paper =
Chemical equivalent of zinc =
Weight of filter paper =
Weight of copper =
Chemical equivalent of copper =

Cut off a piece of pure zinc rod which will weigh not over four grams. Prepare a solution of copper sulphate by dissolving 25 grams of copper sulphate in 200 grams of water held in a beaker. Weigh the zinc accurately (to one centigram if possible). Weigh a filter paper and put it carefully away. Drop the zinc into the beaker containing the copper sulphate solution and cover the beaker with wet filter paper. Leave it until the zinc seems to have wholly disappeared. This will take at least until your next laboratory period. When the substitution seems to be complete, filter the solution into the weighed filter paper. Let the paper and its contents dry for at least a day in an evaporating dish nearly, but not wholly, covered by a watch glass. When dry, weigh the copper which has come out of solution.

B

What is the relation between the weight of the zinc and the weight of the copper?

If the chemical equivalent of zinc (its atomic weight divided by its valence) is 32.7, what do you find to be the chemical equivalent of the copper?

Solve by the use of this proportion:

Weight of zinc used: Weight of copper obtained:: Chemical equivalent of zinc, 32.7: Chemical equivalent of copper.

Exercise 80. Bending Glass.

Apparatus — Burner, flame spreader, glass tubing, thistle tube, asbestos.

Take four pieces of glass tubing about a foot long. With these you are to form right bends, tips, and delivery tubes similar to those shown in the illustration given of a wash bottle Ex. 59. and of the picture accompanying the exercise on hydrogen, Ex. 8. You are also to form an oblique bend delivery tube similar to that shown in the exercise on oxygen, Ex. 5.

Cut two of the pieces in two by first making a slight scratch on one side of the middle of the tube with a three-cornered file, then placing your two thumbs just back of the scratch on the opposite side of the tube and pressing gently. The tube breaks where it was scratched.

Five things must be remembered in bending glass:

1. Use a flame spreader and heat as much of the length of the glass as possible.
2. Use the luminous flame for bending and the non-luminous for drawing out the tips and rounding edges.
3. Keep the glass turning constantly in your hands.
4. Always make a final bend outside the flame.
5. Never lay the glass down until it is cool.

Make one right bend with one of the short pieces of glass, and submit the right bend before going further. Then make two right bends from small pieces. Polish off the rough ends by revolving them in the non-luminous flame. Make two tips from large pieces by heating the center of the tube until soft and then drawing out the ends. Make one oblique bend from a large piece. Let the pieces cool on asbestos, clean them off with a piece of filter paper and put away.

LIST OF GENERAL APPARATUS

There should be a sufficient amount of balances to supply one laboratory section.

Small, German, horn-pan balances make a good balance for such experiments as those given here. If they are furnished in sufficient quantity to supply one full section, they can be used by all sections. Each balance should have with it a set of weights running from one centigram to fifty grams. These weights should be covered with a hinged wooden cover to prevent the annoying, and otherwise inevitable, loss of the smaller weights. Besides this supply of horn-pan balances, one pair of trip scales with two sets of large weights, one English and one metric, should be on hand.

Each laboratory should have as many Bunsen burners with sufficient tubing, and as many ring stands, as there are pupils who use the desks at one time. There should be half a dozen extra ring stands for any cases where more than one liquid must be kept boiling simultaneously. Each sink should have one large sponge. Each table should be provided with a box of safety matches. A number of wooden blocks 2½″ by 2½″ by 1″ should be on hand to support apparatus. These may be cut out by any carpenter or bought of an apparatus house. In addition to these things the laboratory should possess a sufficient number of blow-pipes and graduates of 25 cc. capacity to supply a single section, two 500 cc. graduates, the usual stock of reagents, a few iron plates, sand baths, iron mortars, copper wire, cork borers, and a cork press.

The following apparatus should be in the hands of each student of a laboratory section:

1 Erlenmeyer flask with 2-hole stopper.
6 Test tubes.
2 Ignition tubes, with 2 feet large rubber connector.
1 Funnel.
1 Slow match.
2 Wide-mouth bottles, 4 oz.
1 Pneumatic trough (cake tin).
1 Beaker.
1 Test-tube rack (may be fixed).
1 Crucible with cover.
1 Evaporating dish.
1 Asbestos centered wire gauze.
1 Test-tube holder.
1 Pair forceps.
1 Deflagrating spoon.
1 Mortar and pestle.
1 3-cornered file.
1 Pipe clay triangle.
2 Rubber connectors.
1 Flame spreader.
1 Thistle tube.
5 Pieces glass tubing.
1 Medicine dropper.
1 Block weighing paper
1 Block litmus paper.
1 Paper 20 cm. rule.
1 Watch glass.

Three types of apparatus mentioned above should have some further explanation. Small ignition tubes of very hard glass, approximately 4″ long by ½″ in inside diameter, are valuable for many experiments. These tubes may be closed by small corks, or they may be used without stoppers when pieces of large rubber tubing are slipped directly over their mouth. A small cake tin, 4½″ by 7½″ by 2½″ approximately, makes a very serviceable pneumatic trough. Weighing paper may be simply the ordinary brown practice paper furnished to schools. This can be cut in any printing office to a convenient size, say 3½″ by 4″. When creased in the middle it can and should be used for the transference of most solids used here. Without such paper some contamination of the chemicals is almost inevitable. The apparatus and chemicals mentioned in this book may be purchased from any dealer in apparatus.

The following list gives the names of the chemicals used in this course. The double column is arranged approximately in the order in which each substance first appears and should be read down the column. Copper filings is used first, flowers of sulphur second, lead nitrate third, salt fourth, mercuric oxide fifth, etc., etc. As the amounts to be ordered vary so greatly with conditions, exact quantities have not been written down. The usual unit by which each substance is measured for ordinary laboratory use, whether pounds, ounces, quarts or quires, follows each heading. Every substance mentioned in this book is included in the list, whether needed for student use or not.

PART I

Sulphuric acid, com'l (9 lb. bottles). Small quantity C. P. lbs.
Nitric acid, com'l (7 lb. bottles). Small quantity C. P. lbs.
Hydrochloric acid, com'l (6 lb. bottles). Small quantity C. P. lbs.
Ammonia (4 lb. bottles).

Copper filings, lbs.
Sulphur, lbs.
Lead nitrate, oz.
Sodium chloride (common salt), lbs.
Mercuric oxide, oz.
Slow match, pieces.
Potassium chlorate, lbs.
Manganese dioxide, lbs.
Powdered charcoal, lbs.
Red phosphorus, oz.
Iron watch springs, oz.
Granulated zinc, lbs.
Iron wire, No. 24, oz.
    (Soft iron, cut for dissolving)
Methyl orange, oz.
Phenolphthalein, oz.
Powdered quartz, lbs.
Calcium fluoride, lbs.
Sugar, lbs.
Iron sulphide, lbs.
Potassium hydroxide, lbs.
Sulphurous acid, lbs.
Sodium sulphite, lbs.
Potassium sulphite, lbs.
Lampblack, lbs.
Charcoal blocks, doz.
Red lead, minium, lbs.
Sodium acetate, oz.
Sodium biborate, borax, lbs.
Alcohol (ethyl), qts.
Kerosene, qts.
Molasses, qts.
Acetic acid, qts.

Red litmus paper, quires.
Blue litmus paper, quires.
Turmeric paper, quires.
Vinegar, qts.
Sodium sulphate, lbs.
Sodium hydroxide, lbs.
Sodium nitrate, lbs.
Zinc oxide, oz.
Barium chloride C. P., lbs.
Lithium nitrate, oz.
Calcium nitrate, oz.
Potassium nitrate, oz.
Barium nitrate, oz.
Strontium nitrate, oz.
Silver nitrate, oz.
Iodine crystals, oz.
Starch, lbs.
Potassium iodide, oz.
Potassium bromide, oz.
Carbon bisulphide, lbs.
Calcium carbonate, powdered marble, lbs.
Marble chips, lbs.
Sodium carbonate, lbs.
Calcium oxide, lime, lbs.
Ammonium chloride, lbs.
Sodium nitrite, lbs.
Potassium carbonate, lbs.
Ferrous sulphate, lbs.
Yellow phosphorus, oz.
Copper sulphate, lbs.
Calcium chloride, anhydrous, lbs.
Mercury, lbs.
Sodium silicate, water glass, lbs.
Ethel acetate, oz.
Wood alcohol (methyl), qts.
Castor oil, qts.
Picric acid, oz.
Alizarin paste, oz.
Iron acetate, oz.
Ammonium citrate of iron, oz.
Potassium ferricyanide, oz.
Potassium ferrocyanide, oz.
Ferrous chloride, oz.
Ferric chloride, oz.
Ferrous sulphate, lbs.
Lead wire, oz.
Lead oxide, lbs.
Mercurous nitrate, oz.
Iron, reduced by alcohol, lbs.
Magnesium ribbon, oz.
Sodium amalgam, oz.
Paraffin, oz.
Ammonium carbonate, lbs.
Sodium bicarbonate, lbs.
Potassium chromate, oz.
Potassium bichromate, oz.
Chrome alum, oz.
Cobalt nitrate, oz.
Nickel nitrate, oz.
Zinc wire, C. P., oz.
Glass wool, oz.
Copper oxide, black, oz.
Zinc rod, C. P., oz.
Wood ashes, lbs.
Sand, lbs.
Platinum wire, oz.
Candles, lbs.

Lime water, the solution of calcium hydroxide, should be prepared by filling the bottom of a large bottle with lime and adding water to be renewed as used. This makes a saturated solution from which the amounts required for daily use should be siphoned off to keep the liquid clear. Chlorine water, mentioned in one or two experiments, may be prepared by passing chlorine gas into water. Phenolphthalein should be used in the form of an alcoholic solution.

The following apparatus is needed for a few of the preceding experiments. No numbers are given as to quantities, as some of these things may vary from the single piece necessary for lecture-table use to the number necessary for the equipment of every pupil in the laboratory.

APPARATUS

Barometer,

Boyle's Law tube, with adjustable arm.

Charles' Law tube, straight pattern, mounted on rod.

Burettes, without ground glass stopcocks, using glass balls instead of pinchcocks.

Liebig condensers.

Retorts.

Galvanoscopes.

Batteries.

ATOMIC WEIGHTS AND SYMBOLS OF IMPORTANT ELEMENTS

Name	Symbol	Atomic Weight	
		Accurate O = 16	Approximate (to be used for all calculations)
Aluminum	Al.	27.1	27
Antimony	Sb.	120.2	120
Argon	A.	39.9	
Arsenic	As.	75.0	75
Barium	Ba.	137.37	137
Bismuth	Bi.	208.0	208
Bromine	Br.	79.92	80
Boron	B.	11.0	11
Cadmium	Cd.	112.40	112
Calcium	Ca.	40.09	40
Carbon	C.	12.00	12
Chlorine	Cl.	35.46	35.5
Chromium	Cr.	52.1	52
Cobalt	Co.	58.97	59
Copper	Cu.	63.57	63.5
Fluorine	F.	19.0	19
Gold	Au.	197.2	197
Hydrogen	H.	1.008	1
Iodine	I.	126.92	127
Iron	Fe.	55.85	56
Lead	Pb.	207.10	207
Lithium	Li.	7.00	7
Magnesium	Mg.	24.32	24
Manganese	Mn.	54.93	55
Mercury	Hg.	200.0	200
Nickel	Ni.	58.68	58.5
Nitrogen	N.	14.01	14
Oxygen	O.	16.00	16
Phosphorus	P.	31.0	31
Platinum	Pt.	195.0	195
Potassium	K.	39.10	39
Silicon	Si.	28.3	28
Silver	Ag.	107.88	108
Sodium	Na.	23.00	23
Strontium	Sr.	87.62	87.5
Sulphur	S.	32.07	32
Tellurium	Te.	127.5	
Thorium	Th.	232.42	
Tin	Sn.	119.0	119
Tungsten	W.	184.0	
Zinc	Zn.	65.37	65.4